14

Acids and Bases

A 30-YEAR-OLD MAN HAS been brought to the emergency room after an automobile accident. The emergency room nurses are tending to the patient, Larry, who is unresponsive. A blood sample is taken then sent to Brianna, a clinical laboratory technician, who begins the process of analyzing the pH, the partial pressures of O_2 and CO_2, and the concentrations of glucose and electrolytes.

Within minutes, Brianna determines that Larry's blood pH is 7.30 and the partial pressure of CO_2 gas is above the desired level. Blood pH is typically in the range of 7.35 to 7.45, and a value less than 7.35 indicates a state of acidosis. Respiratory acidosis occurs because an increase in the partial pressure of CO_2 gas in the bloodstream prevents the biochemical buffers in blood from making a change in the pH.

Brianna recognizes these signs and immediately contacts the emergency room to inform them that Larry's airway may be blocked. In the emergency room, they provide Larry with an IV containing bicarbonate to increase the blood pH and begin the process of unblocking his airway. Shortly afterward, Larry's airway is cleared, and his blood pH and partial pressure of CO_2 gas return to normal.

CAREER

Clinical Laboratory Technician

Clinical laboratory technicians, also known as medical laboratory technicians, perform a wide variety of tests on body fluids and cells that help in the diagnosis and treatment of patients. These tests range from determining blood concentrations of glucose and cholesterol to determining drug levels in the blood for transplant patients or a patient undergoing treatment. Clinical laboratory technicians also prepare specimens in the detection of cancerous tumors, and type blood samples for transfusions. Clinical laboratory technicians must also interpret and analyze the test results, which are then passed on to the physician.

CHAPTER READINESS*

KEY MATH SKILLS

- Solving Equations (1.4)
- Converting between Standard Numbers and Scientific Notation (1.5)

CORE CHEMISTRY SKILLS

- Writing Ionic Formulas (6.2)
- Balancing a Chemical Equation (8.2)
- Using Concentration as a Conversion Factor (12.4)
- Writing the Equilibrium Expression (13.3)
- Calculating Equilibrium Concentrations (13.4)
- Using Le Châtelier's Principle (13.5)

*These Key Math Skills and Core Chemistry Skills from previous chapters are listed here for your review as you proceed to the new material in this chapter.

LOOKING AHEAD

14.1 Acids and Bases

LEARNING GOAL Describe and name acids and bases.

Citrus fruits are sour because of the presence of acids.

Acids and bases are important substances in health, industry, and the environment. One of the most common characteristics of acids is their sour taste. Lemons and grapefruits taste sour because they contain acids such as citric and ascorbic acid (vitamin C). Vinegar tastes sour because it contains acetic acid. We produce lactic acid in our muscles when we exercise. Acid from bacteria turns milk sour in the production of yogurt and cottage cheese. We have hydrochloric acid in our stomachs that helps us digest food. Sometimes we take antacids, which are bases such as sodium bicarbonate or milk of magnesia, to neutralize the effects of too much stomach acid.

The term *acid* comes from the Latin word *acidus*, which means "sour." We are familiar with the sour tastes of vinegar and lemons and other common acids in foods.

In 1887, the Swedish chemist Svante Arrhenius was the first to describe **acids** as substances that produce hydrogen ions (H^+) when they dissolve in water. Because acids produce ions in water, they are also electrolytes. For example, hydrogen chloride dissociates in water to give hydrogen ions, H^+, and chloride ions, Cl^-. The hydrogen ions give acids a sour taste, change the blue litmus indicator to red, and corrode some metals.

$$\underset{\text{Polar molecular compound}}{\mathbf{H}\mathrm{Cl}(g)} \xrightarrow[\text{Dissociation}]{H_2O} \underset{\text{Hydrogen ion}}{\mathbf{H}^+(aq)} + Cl^-(aq)$$

Naming Acids

Acids dissolve in water to produce hydrogen ions, along with a negative ion that may be a simple nonmetal anion or a polyatomic ion. When an acid dissolves in water to produce a hydrogen ion and a simple nonmetal anion, the prefix *hydro* is used before the name of the nonmetal, and its *ide* ending is changed to *ic acid*. For example, hydrogen chloride (HCl) dissolves in water to form HCl(*aq*), which is named hydrochloric acid. An exception is hydrogen cyanide (HCN), which as an acid is named hydrocyanic acid.

When an acid contains oxygen, it dissolves in water to produce a hydrogen ion and an oxygen-containing polyatomic anion. The most common form of an oxygen-containing acid has a name that ends with *ic acid*. The name of its polyatomic anion ends in *ate*. If the acid contains a polyatomic ion with an *ite* ending, its name ends in *ous acid*.

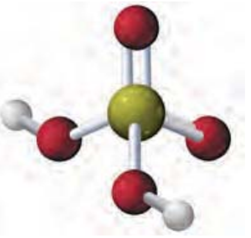
The halogens in Group 7A (17) can form more than two oxygen-containing acids. For chlorine, the common form is chloric acid ($HClO_3$), which contains the chlorate polyatomic ion (ClO_3^-). For the acid that contains one more oxygen atom than the common form, the prefix *per* is used; $HClO_4$ is named *perchloric acid*. When the polyatomic ion in the acid has one oxygen atom less than the common form, the suffix *ous* is used. Thus, $HClO_2$ is named *chlorous acid*; it contains the chlorite ion (ClO_2^-). The prefix *hypo* is used for the acid that has two oxygen atoms less than the common form; HClO is named *hypochlorous acid*. The names of some common acids and their anions are listed in **TABLE 14.1**.

ENGAGE

Why is HBr named hydrobromic acid but $HBrO_3$ is named bromic acid?

TABLE 14.1 Names of Common Acids and Their Anions

Acid	Name of Acid	Anion	Name of Anion
HCl	**Hydro**chlor**ic acid**	Cl^-	Chlor**ide**
HBr	**Hydro**brom**ic acid**	Br^-	Brom**ide**
HI	**Hydro**iod**ic acid**	I^-	Iod**ide**
HCN	**Hydro**cyan**ic acid**	CN^-	Cyan**ide**
HNO_3	Nitr**ic acid**	NO_3^-	Nitr**ate**
HNO_2	Nitr**ous acid**	NO_2^-	Nitr**ite**
H_2SO_4	Sulfur**ic acid**	SO_4^{2-}	Sulf**ate**
H_2SO_3	Sulfur**ous acid**	SO_3^{2-}	Sulf**ite**
H_2CO_3	Carbon**ic acid**	CO_3^{2-}	Carbon**ate**
$HC_2H_3O_2$	Acet**ic acid**	$C_2H_3O_2^-$	Acet**ate**
H_3PO_4	Phosphor**ic acid**	PO_4^{3-}	Phosph**ate**
H_3PO_3	Phosphor**ous acid**	PO_3^{3-}	Phosph**ite**
$HClO_3$	Chlor**ic acid**	ClO_3^-	Chlor**ate**
$HClO_2$	Chlor**ous acid**	ClO_2^-	Chlor**ite**



Sulfuric acid dissolves in water to produce one or two H^+ and an anion.

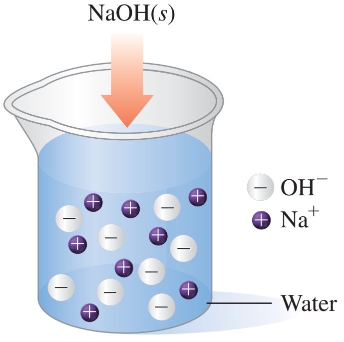
Bases

You may be familiar with some household bases such as antacids, drain openers, and oven cleaners. According to the Arrhenius theory, **bases** are ionic compounds that dissociate into cations and hydroxide ions (OH^-) when they dissolve in water. They are another example of strong electrolytes. For example, sodium hydroxide is an Arrhenius base that dissociates completely in water to give sodium ions (Na^+) and hydroxide ions (OH^-).

Most Arrhenius bases are formed from Groups 1A (1) and 2A (2) metals, such as NaOH, KOH, LiOH, and $Ca(OH)_2$. The hydroxide ions (OH^-) give Arrhenius bases common characteristics, such as a bitter taste and a slippery feel. A base turns litmus indicator blue and phenolphthalein indicator pink. **TABLE 14.2** compares some characteristics of acids and bases.

TABLE 14.2 Some Characteristics of Acids and Bases

Characteristic	Acids	Bases
Arrhenius	Produce H^+	Produce OH^-
Electrolytes	Yes	Yes
Taste	Sour	Bitter, chalky
Feel	May sting	Soapy, slippery
Litmus	Red	Blue
Phenolphthalein	Colorless	Pink
Neutralization	Neutralize bases	Neutralize acids



$$NaOH(s) \xrightarrow{H_2O} Na^+(aq) + OH^-(aq)$$

Ionic compound Dissociation Hydroxide ion

An Arrhenius base produces cations and OH^- anions in an aqueous solution.

Calcium hydroxide, $Ca(OH)_2$, is used in the food industry to produce beverages, and in dentistry as a filler for root canals.

Naming Bases

Typical Arrhenius bases are named as *hydroxides*.

Base	Name
LiOH	Lithium **hydroxide**
NaOH	Sodium **hydroxide**
KOH	Potassium **hydroxide**
$Ca(OH)_2$	Calcium **hydroxide**
$Al(OH)_3$	Aluminum **hydroxide**

A soft drink contains H_3PO_4 and H_2CO_3.

SAMPLE PROBLEM 14.1 Names and Formulas of Acids and Bases

a. Identify each of the following as an acid or a base and give its name:
  1. H_3PO_4, ingredient in soft drinks
  2. NaOH, ingredient in oven cleaner
b. Write the formula for each of the following:
  1. magnesium hydroxide, ingredient in antacids
  2. hydrobromic acid, used industrially to prepare bromide compounds

TRY IT FIRST

SOLUTION

a. 1. acid, phosphoric acid **2.** base, sodium hydroxide
b. 1. $Mg(OH)_2$ **2.** HBr

STUDY CHECK 14.1

a. Identify as an acid or a base and give the name for H_2CO_3.
b. Write the formula for iron(III) hydroxide.

ANSWER

a. acid, carbonic acid **b.** $Fe(OH)_3$

QUESTIONS AND PROBLEMS

14.1 Acids and Bases

LEARNING GOAL Describe and name acids and bases.

14.1 Indicate whether each of the following statements is characteristic of an acid, a base, or both:
a. has a sour taste
b. neutralizes bases
c. produces H^+ ions in water
d. is named barium hydroxide
e. is an electrolyte

14.2 Indicate whether each of the following statements is characteristic of an acid, a base, or both:
a. neutralizes acids
b. produces OH^- ions in water
c. has a slippery feel
d. conducts an electrical current in solution
e. turns litmus red

14.3 Name each of the following acids or bases:
a. HCl **b.** $Ca(OH)_2$ **c.** $HClO_4$
d. HNO_3 **e.** H_2SO_3 **f.** $HBrO_2$

14.4 Name each of the following acids or bases:
a. $Al(OH)_3$ **b.** HBr **c.** H_2SO_4
d. KOH **e.** HNO_2 **f.** $HClO_2$

14.5 Write formulas for each of the following acids and bases:
a. rubidium hydroxide **b.** hydrofluoric acid
c. phosphoric acid **d.** lithium hydroxide
e. ammonium hydroxide **f.** periodic acid

14.6 Write formulas for each of the following acids and bases:
a. barium hydroxide **b.** hydroiodic acid
c. nitric acid **d.** strontium hydroxide
e. acetic acid **f.** hypochlorous acid

14.2 Brønsted–Lowry Acids and Bases

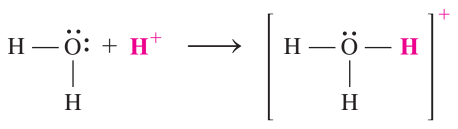
LEARNING GOAL Identify conjugate acid–base pairs for Brønsted–Lowry acids and bases.

In 1923, J. N. Brønsted in Denmark and T. M. Lowry in Great Britain expanded the definition of acids and bases to include bases that do not contain OH^- ions. A **Brønsted–Lowry acid** can donate a hydrogen ion, H^+, and a **Brønsted–Lowry base** can accept a hydrogen ion.

A Brønsted–Lowry acid is a substance that donates H^+.

A Brønsted–Lowry base is a substance that accepts H^+.

A free hydrogen ion does not actually exist in water. Its attraction to polar water molecules is so strong that the H^+ bonds to a water molecule and forms a **hydronium ion, H_3O^+**.

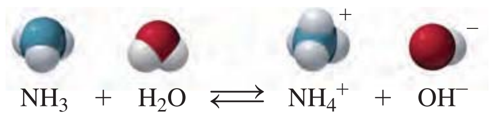


We can write the formation of a hydrochloric acid solution as a transfer of H^+ from hydrogen chloride to water. By accepting an H^+ in the reaction, water is acting as a base according to the Brønsted–Lowry concept.

$$HCl + H_2O \longrightarrow H_3O^+ + Cl^-$$

Hydrogen chloride — Acid (H^+ donor); Water — Base (H^+ acceptor); Hydronium ion; Chloride ion — Acidic solution

In another reaction, ammonia (NH_3) acts as a base by accepting H^+ when it reacts with water. Because the nitrogen atom of NH_3 has a stronger attraction for H^+ than oxygen, water acts as an acid by donating H^+.



$$NH_3 + H_2O \rightleftarrows NH_4^+ + OH^-$$

Ammonia — Base (H^+ acceptor); Water — Acid (H^+ donor); Ammonium ion; Hydroxide ion — Basic solution

SAMPLE PROBLEM 14.2 Acids and Bases

In each of the following equations, identify the reactant that is a Brønsted–Lowry acid and the reactant that is a Brønsted–Lowry base:

a. $HBr(aq) + H_2O(l) \longrightarrow H_3O^+(aq) + Br^-(aq)$
b. $CN^-(aq) + H_2O(l) \rightleftarrows HCN(aq) + OH^-(aq)$

TRY IT FIRST

SOLUTION

a. HBr, Brønsted–Lowry acid; H_2O, Brønsted–Lowry base
b. H_2O, Brønsted–Lowry acid; CN^-, Brønsted–Lowry base

STUDY CHECK 14.2

When HNO_3 reacts with water, water acts as a Brønsted–Lowry base. Write the equation for the reaction.

ANSWER

$$HNO_3(aq) + H_2O(l) \longrightarrow H_3O^+(aq) + NO_3^-(aq)$$

CORE CHEMISTRY SKILL
Identifying Conjugate Acid–Base Pairs

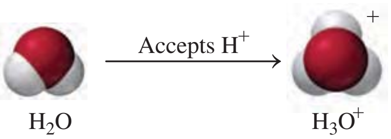
Conjugate Acid–Base Pairs

According to the Brønsted–Lowry theory, a **conjugate acid–base pair** consists of molecules or ions related by the loss of one H^+ by an acid, and the gain of one H^+ by a base. Every acid–base reaction contains two conjugate acid–base pairs because an H^+ is transferred in both the forward and reverse directions. When an acid such as HF loses one H^+, the conjugate base F^- is formed. When the base H_2O gains an H^+, its conjugate acid, H_3O^+, is formed.

Because the overall reaction of HF is reversible, the conjugate acid H_3O^+ can donate H^+ to the conjugate base F^- and re-form the acid HF and the base H_2O. Using the relationship of loss and gain of one H^+, we can now identify the conjugate acid–base pairs as HF/F^- along with H_3O^+/H_2O.

Conjugate acid–base pair

Conjugate acid–base pair



$$\underset{\text{Acid}}{HF(aq)} + \underset{\text{Base}}{H_2O(l)} \rightleftharpoons \underset{\text{Conjugate base}}{F^-(aq)} + \underset{\text{Conjugate acid}}{H_3O^+(aq)}$$

H^+ gain
H^+ loss
H^+ gain
H^+ loss

HF, an acid, loses one H^+ to form its conjugate base F^-. Water acts as a base by gaining one H^+ to form its conjugate acid H_3O^+.

ENGAGE
Why is $HBrO_2$ the conjugate acid of BrO_2^-?

In another reaction, ammonia (NH_3) accepts H^+ from H_2O to form the conjugate acid NH_4^+ and conjugate base OH^-. Each of these conjugate acid–base pairs, NH_4^+/NH_3 and H_2O/OH^-, is related by the loss and gain of one H^+.

Conjugate acid–base pair

$$NH_3(g) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$$

Conjugate acid–base pair

Ammonia, NH_3, acts as a base when it gains one H^+ to form its conjugate acid NH_4^+. Water acts as an acid by losing one H^+ to form its conjugate base OH^-.

In these two examples, we see that water can act as an acid when it donates H^+ or as a base when it accepts H^+. Substances that can act as both acids and bases are **amphoteric** or *amphiprotic*. For water, the most common amphoteric substance, the acidic or basic behavior depends on the other reactant. Water donates H^+ when it reacts with a stronger base, and it accepts H^+ when it reacts with a stronger acid. Another example of an amphoteric substance is bicarbonate (HCO_3^-). With a base, HCO_3^- acts as an acid and donates one H^+ to give CO_3^{2-}. However, when HCO_3^- reacts with an acid, it acts as a base and accepts one H^+ to form H_2CO_3.

ENGAGE
Why can H_2O be both the conjugate base of H_3O^+ and the conjugate acid of OH^-?

Amphoteric substances act as both acids and bases.

H_3O^+ / H_2CO_3	← Acts as a base	H_2O / HCO_3^-	Acts as an acid →	OH^- / CO_3^{2-}

SAMPLE PROBLEM 14.3 Identifying Conjugate Acid–Base Pairs

Identify the conjugate acid–base pairs in the following reaction:

$$HBr(aq) + NH_3(aq) \longrightarrow Br^-(aq) + NH_4^+(aq)$$

TRY IT FIRST

SOLUTION

ANALYZE THE PROBLEM	Given		Need	Connect
	HBr NH_3	Br^- NH_4^+	conjugate acid–base pairs	lose/gain one H^+

STEP 1 Identify the reactant that loses H^+ as the acid. In the reaction, HBr donates H^+ to form the product Br^-. Thus HBr is the acid and Br^- is its conjugate base.

STEP 2 Identify the reactant that gains H^+ as the base. In the reaction, NH_3 gains H^+ to form the product NH_4^+. Thus, NH_3 is the base and NH_4^+ is its conjugate acid.

STEP 3 Write the conjugate acid–base pairs.

HBr/Br^- and NH_4^+/NH_3

STUDY CHECK 14.3

Identify the conjugate acid–base pairs in the following reaction:

$$HCN(aq) + SO_4^{2-}(aq) \rightleftharpoons CN^-(aq) + HSO_4^-(aq)$$

ANSWER

The conjugate acid–base pairs are HCN/CN^- and HSO_4^-/SO_4^{2-}.

Guide to Writing Conjugate Acid–Base Pairs

STEP 1 Identify the reactant that loses H^+ as the acid.

STEP 2 Identify the reactant that gains H^+ as the base.

STEP 3 Write the conjugate acid–base pairs.

QUESTIONS AND PROBLEMS

14.2 Brønsted–Lowry Acids and Bases

LEARNING GOAL Identify conjugate acid–base pairs for Brønsted–Lowry acids and bases.

14.7 Identify the reactant that is a Brønsted–Lowry acid and the reactant that is a Brønsted–Lowry base in each of the following:

a. $HI(aq) + H_2O(l) \longrightarrow I^-(aq) + H_3O^+(aq)$
b. $F^-(aq) + H_2O(l) \rightleftharpoons HF(aq) + OH^-(aq)$
c. $H_2S(aq) + CH_3-CH_2-NH_2(aq) \rightleftharpoons HS^-(aq) + CH_3-CH_2-NH_3^+(aq)$

14.8 Identify the reactant that is a Brønsted–Lowry acid and the reactant that is a Brønsted–Lowry base in each of the following:

a. $CO_3^{2-}(aq) + H_2O(l) \rightleftharpoons HCO_3^-(aq) + OH^-(aq)$
b. $H_2SO_4(aq) + H_2O(l) \longrightarrow HSO_4^-(aq) + H_3O^+(aq)$
c. $C_2H_3O_2^-(aq) + H_3O^+(aq) \rightleftharpoons HC_2H_3O_2(aq) + H_2O(l)$

14.9 Write the formula for the conjugate base for each of the following acids:

a. HF b. H_2O c. $H_2PO_3^-$
d. HSO_4^- e. $HClO_2$

14.10 Write the formula for the conjugate base for each of the following acids:

a. HCO_3^- b. $CH_3-NH_3^+$ c. HPO_4^{2-}
d. HNO_2 e. HBrO

14.11 Write the formula for the conjugate acid for each of the following bases:

a. CO_3^{2-} b. H_2O c. $H_2PO_4^-$
d. Br^- e. ClO_4^-

14.12 Write the formula for the conjugate acid for each of the following bases:

a. SO_4^{2-} b. CN^- c. NH_3
d. ClO_2^- e. HS^-

14.13 Identify the Brønsted–Lowry acid–base pairs in each of the following equations:

a. $H_2CO_3(aq) + H_2O(l) \rightleftharpoons HCO_3^-(aq) + H_3O^+(aq)$
b. $NH_4^+(aq) + H_2O(l) \rightleftharpoons NH_3(aq) + H_3O^+(aq)$
c. $HCN(aq) + NO_2^-(aq) \rightleftharpoons CN^-(aq) + HNO_2(aq)$
d. $CHO_2^-(aq) + HF(aq) \rightleftharpoons HCHO_2(aq) + F^-(aq)$

14.14 Identify the Brønsted–Lowry acid–base pairs in each of the following equations:

a. $H_3PO_4(aq) + H_2O(l) \rightleftharpoons H_2PO_4^-(aq) + H_3O^+(aq)$
b. $CO_3^{2-}(aq) + H_2O(l) \rightleftharpoons HCO_3^-(aq) + OH^-(aq)$
c. $H_3PO_4(aq) + NH_3(aq) \rightleftharpoons H_2PO_4^-(aq) + NH_4^+(aq)$
d. $HNO_2(aq) + CH_3-CH_2-NH_2(aq) \rightleftharpoons CH_3-CH_2-NH_3^+(aq) + NO_2^-(aq)$

14.15 When ammonium chloride dissolves in water, the ammonium ion NH_4^+ donates an H^+ to water. Write a balanced equation for the reaction of the ammonium ion with water.

14.16 When sodium carbonate dissolves in water, the carbonate ion CO_3^{2-} acts as a base. Write a balanced equation for the reaction of the carbonate ion with water.

14.3 Strengths of Acids and Bases

LEARNING GOAL Write equations for the dissociation of strong and weak acids; identify the direction of reaction.

In the process called **dissociation**, an acid or a base separates into ions in water. The *strength* of an acid is determined by the moles of H_3O^+ that are produced for each mole of acid that dissociates. The *strength* of a base is determined by the moles of OH^- that are produced for each mole of base that dissolves. Strong acids and strong bases dissociate completely in water, whereas weak acids and weak bases dissociate only slightly, leaving most of the initial acid or base undissociated.

Weak acids are found in foods and household products.

Strong and Weak Acids

Strong acids are examples of strong electrolytes because they donate H^+ so easily that their dissociation in water is essentially complete. For example, when HCl, a strong acid, dissociates in water, H^+ is transferred to H_2O; the resulting solution contains essentially only the ions H_3O^+ and Cl^-. We consider the reaction of HCl in H_2O as going 100% to products. Thus, one mole of a strong acid dissociates in water to yield one mole of H_3O^+ and one mole of its conjugate base. We write the equation for a strong acid such as HCl with a single arrow.

$$HCl(g) + H_2O(l) \longrightarrow H_3O^+(aq) + Cl^-(aq)$$

There are only six common strong acids, which are stronger acids than H_3O^+. All other acids are weak. **TABLE 14.3** lists the relative strengths of acids and bases. **Weak**

TABLE 14.3 Relative Strengths of Acids and Bases

(Acid Strength Increases ↑ ; Base Strength Increases ↓)

Acid		Conjugate Base	
Strong Acids			
Hydroiodic acid	HI	I^-	Iodide ion
Hydrobromic acid	HBr	Br^-	Bromide ion
Perchloric acid	$HClO_4$	ClO_4^-	Perchlorate ion
Hydrochloric acid	HCl	Cl^-	Chloride ion
Sulfuric acid	H_2SO_4	HSO_4^-	Hydrogen sulfate ion
Nitric acid	HNO_3	NO_3^-	Nitrate ion
Hydronium ion	H_3O^+	H_2O	Water
Weak Acids			
Hydrogen sulfate ion	HSO_4^-	SO_4^{2-}	Sulfate ion
Phosphoric acid	H_3PO_4	$H_2PO_4^-$	Dihydrogen phosphate ion
Nitrous acid	HNO_2	NO_2^-	Nitrite ion
Hydrofluoric acid	HF	F^-	Fluoride ion
Acetic acid	$HC_2H_3O_2$	$C_2H_3O_2^-$	Acetate ion
Carbonic acid	H_2CO_3	HCO_3^-	Bicarbonate ion
Hydrosulfuric acid	H_2S	HS^-	Hydrogen sulfide ion
Dihydrogen phosphate ion	$H_2PO_4^-$	HPO_4^{2-}	Hydrogen phosphate ion
Ammonium ion	NH_4^+	NH_3	Ammonia
Hydrocyanic acid	HCN	CN^-	Cyanide ion
Bicarbonate ion	HCO_3^-	CO_3^{2-}	Carbonate ion
Methylammonium ion	$CH_3{-}NH_3^+$	$CH_3{-}NH_2$	Methylamine
Hydrogen phosphate ion	HPO_4^{2-}	PO_4^{3-}	Phosphate ion
Water	H_2O	OH^-	Hydroxide ion

acids are weak electrolytes because they dissociate slightly in water, forming only a small amount of H_3O^+ ions. A weak acid has a strong conjugate base, which is why the reverse reaction is more prevalent. Even at high concentrations, weak acids produce low concentrations of H_3O^+ ions (see **FIGURE 14.1**).

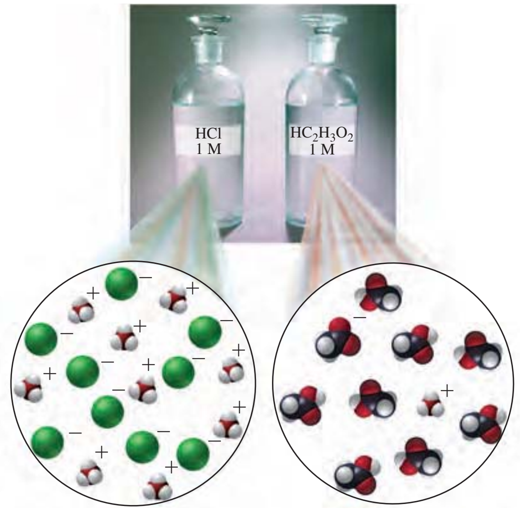


FIGURE 14.1 ▶ A strong acid such as HCl is completely dissociated, whereas a weak acid such as $HC_2H_3O_2$ contains mostly molecules and a few ions.

Q What is the difference between a strong acid and a weak acid?

Many of the products you use at home contain weak acids. Citric acid is a weak acid found in fruits and fruit juices such as lemons, oranges, and grapefruit. The vinegar used in salad dressings is typically a 5% (m/v) acetic acid ($HC_2H_3O_2$) solution. In water, a few $HC_2H_3O_2$ molecules donate H^+ to H_2O to form H_3O^+ ions and acetate ions ($C_2H_3O_2^-$). The reverse reaction also takes place, which converts the H_3O^+ ions and acetate ions ($C_2H_3O_2^-$) back to reactants. The formation of hydronium ions from vinegar is the reason we notice the sour taste of vinegar. We write the equation for a weak acid in an aqueous solution with a double arrow to indicate that the forward and reverse reactions are at equilibrium.

ENGAGE

Why is NO_2^- a stronger base than NO_3^-?

$$\underset{\text{Acetic acid}}{HC_2H_3O_2(aq)} + H_2O(l) \rightleftarrows \underset{\text{Acetate ion}}{C_2H_3O_2^-(aq)} + H_3O^+(aq)$$

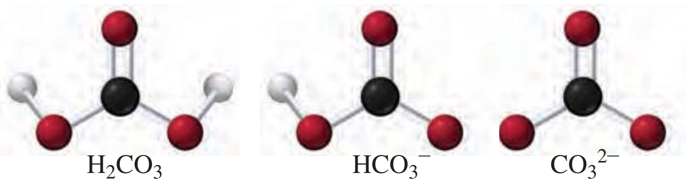
Diprotic Acids

Some weak acids, such as carbonic acid, are *diprotic acids* that have two H^+, which dissociate one at a time. For example, carbonated soft drinks are prepared by dissolving CO_2 in water to form carbonic acid (H_2CO_3). A weak acid such as H_2CO_3 reaches equilibrium between the mostly undissociated H_2CO_3 molecules and the ions H_3O^+ and HCO_3^-.

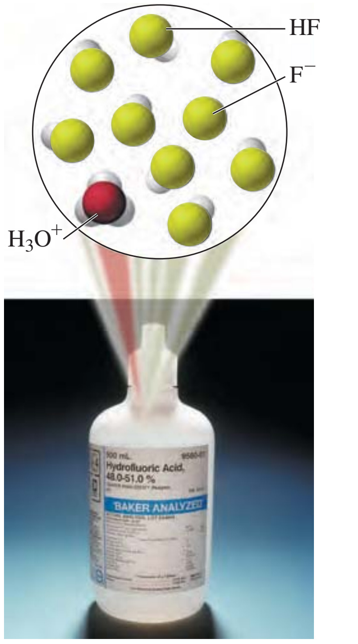
$$\underset{\text{Carbonic acid}}{H_2CO_3(aq)} + H_2O(l) \rightleftarrows H_3O^+(aq) + \underset{\substack{\text{Bicarbonate ion} \\ \text{(hydrogen carbonate)}}}{HCO_3^-(aq)}$$

Because HCO_3^- is also a weak acid, a second dissociation can take place to produce another hydronium ion and the carbonate ion (CO_3^{2-}).

$$\underset{\substack{\text{Bicarbonate ion} \\ \text{(hydrogen carbonate)}}}{HCO_3^-(aq)} + H_2O(l) \rightleftarrows H_3O^+(aq) + \underset{\text{Carbonate ion}}{CO_3^{2-}(aq)}$$



Carbonic acid, a weak acid, loses one H^+ to form hydrogen carbonate ion, which loses a second H^+ to form carbonate ion.



Hydrofluoric acid is the only halogen acid that is a weak acid.

Sulfuric acid (H_2SO_4) is also a diprotic acid. However, its first dissociation is complete (100%), which means H_2SO_4 is a strong acid. The product, hydrogen sulfate (HSO_4^-) can dissociate again but only slightly, which means that the hydrogen sulfate ion is a weak acid.

$$\underset{\text{Sulfuric acid}}{H_2SO_4(aq)} + H_2O(l) \longrightarrow H_3O^+(aq) + \underset{\substack{\text{Bisulfate ion}\\\text{(hydrogen sulfate)}}}{HSO_4^-(aq)}$$

$$\underset{\substack{\text{Bisulfate ion}\\\text{(hydrogen sulfate)}}}{HSO_4^-(aq)} + H_2O(l) \rightleftarrows H_3O^+(aq) + \underset{\text{Sulfate ion}}{SO_4^{2-}(aq)}$$

In summary, a strong acid such as HI in water dissociates completely to form an aqueous solution of the ions H_3O^+ and I^-. A weak acid such as HF dissociates only slightly in water to form an aqueous solution that consists mostly of HF molecules and a few H_3O^+ and F^- ions (see **FIGURE 14.2**).

$$\text{Strong acid:}\quad HI(aq) + H_2O(l) \longrightarrow H_3O^+(aq) + I^-(aq) \quad \text{Completely dissociated}$$
$$\text{Weak acid:}\quad HF(aq) + H_2O(l) \rightleftarrows H_3O^+(aq) + F^-(aq) \quad \text{Slightly dissociated}$$

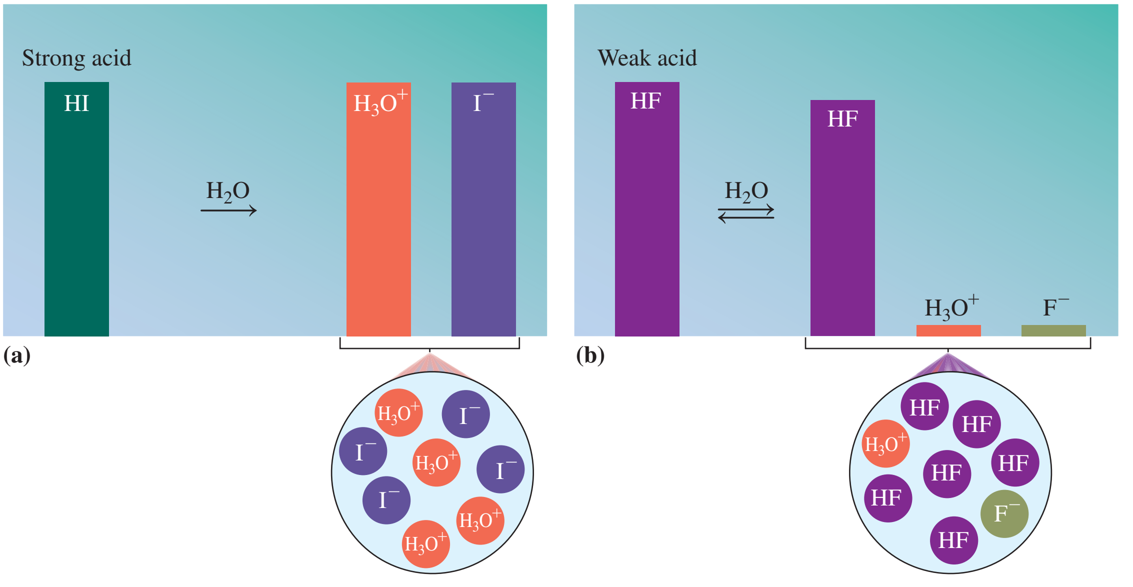


FIGURE 14.2 ▶ After dissociation in water, **(a)** the strong acid HI has high concentrations of H_3O^+ and I^-, and **(b)** the weak acid HF has a high concentration of HF and low concentrations of H_3O^+ and F^-.

Q How do the heights of H_3O^+ and F^- compare to the height of the weak acid HF in the bar diagram for HF?

Strong and Weak Bases

As strong electrolytes, **strong bases** dissociate completely in water. Because these strong bases are ionic compounds, they dissociate in water to give an aqueous solution of metal ions and hydroxide ions. The Group 1A (1) hydroxides are very soluble in water, which can give high concentrations of OH^- ions. A few strong bases are less soluble in water, but what does dissolve dissociates completely as ions. For example, when KOH forms a KOH solution, it contains only the ions K^+ and OH^-.

$$KOH(s) \xrightarrow{H_2O} K^+(aq) + OH^-(aq)$$

Strong Bases

Lithium hydroxide (LiOH)
Sodium hydroxide (NaOH)
Potassium hydroxide (KOH)
Rubidium hydroxide (RbOH)
Cesium hydroxide (CsOH)
Calcium hydroxide ($Ca(OH)_2$)*
Strontium hydroxide ($Sr(OH)_2$)*
Barium hydroxide ($Ba(OH)_2$)*

*Low solubility

Weak bases are weak electrolytes that are poor acceptors of hydrogen ions and produce very few ions in solution. A typical weak base, ammonia (NH_3) is found in window cleaners. In an aqueous solution, only a few ammonia molecules accept hydrogen ions to form NH_4^+ and OH^-.

$$\underset{\text{Ammonia}}{NH_3(g)} + H_2O(l) \rightleftharpoons \underset{\text{Ammonium hydroxide}}{NH_4^+(aq) + OH^-(aq)}$$

Bases in household products are used to remove grease and to open drains.

Bases Used in Household Products

Weak Bases

Window cleaner, ammonia, NH_3
Bleach, NaOCl
Laundry detergent, Na_2CO_3, Na_3PO_4
Toothpaste and baking soda, $NaHCO_3$
Baking powder, scouring powder, Na_2CO_3
Lime for lawns and agriculture, $CaCO_3$
Laxatives, antacids, $Mg(OH)_2$, $Al(OH)_3$

Strong Bases

Drain cleaner, oven cleaner, NaOH

Direction of Reaction

There is a relationship between the components in each conjugate acid–base pair. Strong acids have weak conjugate bases that do not readily accept H^+. As the strength of the acid decreases, the strength of its conjugate base increases.

In any acid–base reaction, there are two acids and two bases. However, one acid is stronger than the other acid, and one base is stronger than the other base. By comparing their relative strengths, we can determine the direction of the reaction. For example, the strong acid H_2SO_4 readily gives up H^+ to water. The hydronium ion H_3O^+ produced is a weaker acid than H_2SO_4, and the conjugate base HSO_4^- is a weaker base than water.

$$\underset{\substack{\text{Stronger}\\\text{acid}}}{H_2SO_4(aq)} + \underset{\substack{\text{Stronger}\\\text{base}}}{H_2O(l)} \longrightarrow \underset{\substack{\text{Weaker}\\\text{acid}}}{H_3O^+(aq)} + \underset{\substack{\text{Weaker}\\\text{base}}}{HSO_4^-(aq)} \quad \text{Mostly products}$$

Let's look at another reaction in which water donates one H^+ to carbonate (CO_3^{2-}) to form HCO_3^- and OH^-. From Table 14.3, we see that HCO_3^- is a stronger acid than H_2O. We also see that OH^- is a stronger base than CO_3^{2-}. To reach equilibrium, the stronger acid and stronger base react in the direction of the weaker acid and weaker base.

$$\underset{\substack{\text{Weaker}\\\text{base}}}{CO_3^{2-}(aq)} + \underset{\substack{\text{Weaker}\\\text{acid}}}{H_2O(l)} \rightleftharpoons \underset{\substack{\text{Stronger}\\\text{acid}}}{HCO_3^-(aq)} + \underset{\substack{\text{Stronger}\\\text{base}}}{OH^-(aq)} \quad \text{Mostly reactants}$$

SAMPLE PROBLEM 14.4 Direction of Reaction

Does the equilibrium mixture of the following reaction contain mostly reactants or products?

$$HF(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + F^-(aq)$$

TRY IT FIRST

SOLUTION

From Table 14.3, we see that HF is a weaker acid than H_3O^+ and that H_2O is a weaker base than F^-. Thus, the equilibrium mixture contains mostly reactants.

$$\underset{\substack{\text{Weaker}\\\text{acid}}}{HF(aq)} + \underset{\substack{\text{Weaker}\\\text{base}}}{H_2O(l)} \rightleftharpoons \underset{\substack{\text{Stronger}\\\text{acid}}}{H_3O^+(aq)} + \underset{\substack{\text{Stronger}\\\text{base}}}{F^-(aq)} \quad \text{Mostly reactants}$$

STUDY CHECK 14.4

Does the equilibrium mixture for the reaction of nitric acid and water contain mostly reactants or products?

ANSWER

$$HNO_3(aq) + H_2O(l) \longrightarrow H_3O^+(aq) + NO_3^-(aq)$$ Mostly products

The equilibrium mixture contains mostly products because HNO_3 is a stronger acid than H_3O^+, and H_2O is a stronger base than NO_3^-.

QUESTIONS AND PROBLEMS

14.3 Strengths of Acids and Bases

LEARNING GOAL Write equations for the dissociation of strong and weak acids; identify the direction of reaction.

14.17 What is meant by the phrase "A strong acid has a weak conjugate base"?

14.18 What is meant by the phrase "A weak acid has a strong conjugate base"?

14.19 Identify the stronger acid in each of the following pairs:
a. HBr or HNO_2 **b.** H_3PO_4 or HSO_4^-
c. HCN or H_2CO_3

14.20 Identify the stronger acid in each of the following pairs:
a. NH_4^+ or H_3O^+ **b.** H_2SO_4 or HCl
c. H_2O or H_2CO_3

14.21 Identify the weaker acid in each of the following pairs:
a. HCl or HSO_4^- **b.** HNO_2 or HF
c. HCO_3^- or NH_4^+

14.22 Identify the weaker acid in each of the following pairs:
a. HNO_3 or HCO_3^- **b.** HSO_4^- or H_2O
c. $H_2SO_4^-$ or H_2CO_3

14.23 Predict whether each of the following reactions contains mostly reactants or products at equilibrium:
a. $H_2CO_3(aq) + H_2O(l) \rightleftharpoons HCO_3^-(aq) + H_3O^+(aq)$
b. $NH_4^+(aq) + H_2O(l) \rightleftharpoons NH_3(aq) + H_3O^+(aq)$
c. $HNO_2(aq) + NH_3(aq) \rightleftharpoons NO_2^-(aq) + NH_4^+(aq)$

14.24 Predict whether each of the following reactions contains mostly reactants or products at equilibrium:
a. $H_3PO_4(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + H_2PO_4^-(aq)$
b. $CO_3^{2-}(aq) + H_2O(l) \rightleftharpoons OH^-(aq) + HCO_3^-(aq)$
c. $HS^-(aq) + F^-(aq) \rightleftharpoons HF(aq) + S^{2-}(aq)$

14.25 Write an equation for the acid–base reaction between ammonium ion and sulfate ion. Why does the equilibrium mixture contain mostly reactants?

14.26 Write an equation for the acid–base reaction between nitrous acid and hydroxide ion. Why does the equilibrium mixture contain mostly products?

14.4 Dissociation Constants for Acids and Bases

LEARNING GOAL Write the dissociation expression for a weak acid or weak base.

As we have seen, acids have different strengths depending on how much they dissociate in water. Because the dissociation of strong acids in water is essentially complete, the reaction is not considered to be an equilibrium situation. However, because weak acids in water dissociate only slightly, the ion products reach equilibrium with the undissociated weak acid molecules. For example, formic acid ($HCHO_2$) the acid found in bee and ant stings, is a weak acid. Formic acid is a weak acid that dissociates in water to form hydronium ion (H_3O^+) and formate ion (CHO_2^-).

$$HCHO_2(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + CHO_2^-(aq)$$

$HCHO_2$

CHO_2^-

Formic acid, a weak acid, loses one H^+ to form formate ion.

Writing Dissociation Constants

An **acid dissociation expression, K_a**, can be written for weak acids that gives the ratio of the concentrations of products to the weak acid reactants. As with other dissociation expressions, the molar concentration of the products is divided by the molar concentration of the reactants. Because water is a pure liquid with a constant concentration, it is omitted. The numerical value of the acid dissociation expression is the acid dissociation constant. For example, the acid dissociation expression for the equilibrium equation of formic acid shown above is written

$$K_a = \frac{[H_3O^+][CHO_2^-]}{[HCHO_2]}$$

The numerical value of the K_a for formic acid at 25 °C is determined by experiment to be 1.8×10^{-4}. Thus, for the weak acid $HCHO_2$, the K_a is written

$$K_a = \frac{[H_3O^+][CHO_2^-]}{[HCHO_2]} = 1.8 \times 10^{-4} \quad \text{Acid dissociation constant}$$

The K_a for formic acid is small, which confirms that the equilibrium mixture of formic acid in water contains mostly reactants and only small amounts of the products. (Recall that the brackets in the K_a represent the molar concentrations of the reactants and products). Weak acids have small K_a values. However, strong acids, which are essentially 100% dissociated, have very large K_a values, but these values are not usually given. **TABLE 14.4** gives K_a and K_b values for selected weak acids and bases.

ENGAGE

Why is an acid with a $K_a = 1.8 \times 10^{-5}$ a stronger acid than an acid with a $K_a = 6.2 \times 10^{-8}$?

TABLE 14.4 K_a and K_b Values for Selected Weak Acids and Bases

Acids		K_a
Phosphoric acid	H_3PO_4	7.5×10^{-3}
Nitrous acid	HNO_2	4.5×10^{-4}
Hydrofluoric acid	HF	3.5×10^{-4}
Formic acid	$HCHO_2$	1.8×10^{-4}
Acetic acid	$HC_2H_3O_2$	1.8×10^{-5}
Carbonic acid	H_2CO_3	4.3×10^{-7}
Hydrosulfuric acid	H_2S	9.1×10^{-8}
Dihydrogen phosphate	$H_2PO_4^-$	6.2×10^{-8}
Hydrocyanic acid	HCN	4.9×10^{-10}
Hydrogen carbonate	HCO_3^-	5.6×10^{-11}
Hydrogen phosphate	HPO_4^{2-}	2.2×10^{-13}
Bases		K_b
Methylamine	$CH_3{-}NH_2$	4.4×10^{-4}
Carbonate	CO_3^{2-}	2.2×10^{-4}
Ammonia	NH_3	1.8×10^{-5}

Let us now consider the dissociation of the weak base methylamine:

$$CH_3{-}NH_2(aq) + H_2O(l) \rightleftarrows CH_3{-}NH_3^+(aq) + OH^-(aq)$$

As we did with the acid dissociation expression, the concentration of water is omitted from the **base dissociation expression, K_b**. The base dissociation constant for methylamine is written

$$K_b = \frac{[CH_3{-}NH_3^+][OH^-]}{[CH_3{-}NH_2]} = 4.4 \times 10^{-4}$$

TABLE 14.5 summarizes the characteristics of acids and bases in terms of strength and equilibrium position.

TABLE 14.5 Characteristics of Acids and Bases

Characteristic	Strong Acids	Weak Acids
Equilibrium Position	Toward products	Toward reactants
K_a	Large	Small
$[H_3O^+]$ and $[A^-]$	100% of [HA] dissociates	Small percent of [HA] dissociates
Conjugate Base	Weak	Strong
Characteristic	**Strong Bases**	**Weak Bases**
Equilibrium Position	Toward products	Toward reactants
K_b	Large	Small
$[BH^+]$ and $[OH^-]$	100% of [B] reacts	Small percent of [B] reacts
Conjugate Acid	Weak	Strong

SAMPLE PROBLEM 14.5 Writing an Acid Dissociation Expression

Write the acid dissociation expression for the weak acid nitrous acid.

TRY IT FIRST

SOLUTION

Guide to Writing the Acid Dissociation Expression

STEP 1 Write the balanced chemical equation.

STEP 2 Write the concentrations of the products as the numerator and the reactants as the denominator.

STEP 1 **Write the balanced chemical equation.** The equation for the dissociation of nitrous acid is written

$$HNO_2(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + NO_2^-(aq)$$

STEP 2 **Write the concentrations of the products as the numerator and the reactants as the denominator.** The acid dissociation expression is written as the concentration of the products divided by the concentration of the undissociated weak acid.

$$K_a = \frac{[H_3O^+][NO_2^-]}{[HNO_2]}$$

STUDY CHECK 14.5

Write the acid dissociation expression for hydrogen phosphate (HPO_4^{2-}).

ANSWER

$$K_a = \frac{[H_3O^+][PO_4^{3-}]}{[HPO_4^{2-}]}$$

QUESTIONS AND PROBLEMS

14.4 Dissociation Constants for Acids and Bases

LEARNING GOAL Write the dissociation expression for a weak acid or weak base.

14.27 Answer *true* or *false* for each of the following: A strong acid
- **a.** is completely dissociated in aqueous solution
- **b.** has a small value of K_a
- **c.** has a strong conjugate base
- **d.** has a weak conjugate base
- **e.** is slightly dissociated in aqueous solution

14.28 Answer *true* or *false* for each of the following: A weak acid
- **a.** is completely dissociated in aqueous solution
- **b.** has a small value of K_a
- **c.** has a strong conjugate base
- **d.** has a weak conjugate base
- **e.** is slightly dissociated in aqueous solution

14.29 Consider the following acids and their dissociation constants:

$$H_2SO_3(aq) + H_2O(l) \rightleftarrows H_3O^+(aq) + HSO_3^-(aq) \quad K_a = 1.2 \times 10^{-2}$$

$$HS^-(aq) + H_2O(l) \rightleftarrows H_3O^+(aq) + S^{2-}(aq) \quad K_a = 1.3 \times 10^{-19}$$

a. Which is the stronger acid, H_2SO_3 or HS^-?
b. What is the conjugate base of H_2SO_3?
c. Which acid has the weaker conjugate base?
d. Which acid has the stronger conjugate base?
e. Which acid produces more ions?

14.30 Consider the following acids and their dissociation constants:

$$HPO_4^{2-}(aq) + H_2O(l) \rightleftarrows H_3O^+(aq) + PO_4^{3-}(aq) \quad K_a = 2.2 \times 10^{-13}$$

$$HCHO_2(aq) + H_2O(l) \rightleftarrows H_3O^+(aq) + CHO_2^-(aq) \quad K_a = 1.8 \times 10^{-4}$$

a. Which is the weaker acid, HPO_4^{2-} or $HCHO_2$?
b. What is the conjugate base of HPO_4^{2-}?
c. Which acid has the weaker conjugate base?
d. Which acid has the stronger conjugate base?
e. Which acid produces more ions?

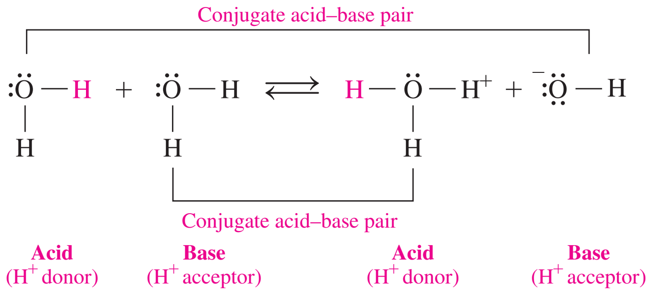
14.31 Phosphoric acid dissociates to form hydronium ion and dihydrogen phosphate. Phosphoric acid has a K_a of 7.5×10^{-3}. Write the equation for the reaction and the acid dissociation expression for phosphoric acid.

14.32 Aniline, $C_6H_5—NH_2$, a weak base with a K_b of 4.0×10^{-10}, reacts with water to form $C_6H_5—NH_3^+$ and hydroxide ion. Write the equation for the reaction and the base dissociation expression for aniline.

14.5 Dissociation of Water

LEARNING GOAL Use the water dissociation expression to calculate the $[H_3O^+]$ and $[OH^-]$ in an aqueous solution.

In many acid–base reactions, water is *amphoteric*, which means that it can act either as an acid or as a base. In pure water, there is a forward reaction between two water molecules that transfers H^+ from one water molecule to the other. One molecule acts as an acid by losing H^+, and the water molecule that gains H^+ acts as a base. Every time H^+ is transferred between two water molecules, the products are one H_3O^+ and one OH^-, which react in the reverse direction to re-form two water molecules. Thus, equilibrium is reached between the conjugate acid–base pairs of water.



Writing the Water Dissociation Expression, K_w

Using the equation for water at equilibrium, we can write its equilibrium expression that shows the concentrations of the products divided by the concentrations of the reactants. Recall that square brackets around the symbols indicate their concentrations in moles per liter (M).

$$H_2O(l) + H_2O(l) \rightleftarrows H_3O^+(aq) + OH^-(aq)$$

$$K = \frac{[H_3O^+][OH^-]}{[H_2O][H_2O]}$$

By omitting the constant concentration of pure water, we can write the **water dissociation expression, K_w.**

$$K_w = [H_3O^+][OH^-]$$

Experiments have determined, that in pure water, the concentration of H_3O^+ and OH^- at 25 °C are each 1.0×10^{-7} M.

Pure water $[H_3O^+] = [OH^-] = 1.0 \times 10^{-7}$ M

When we place the $[H_3O^+]$ and $[OH^-]$ into the water dissociation expression, we obtain the numerical value of K_w, which is 1.0×10^{-14} at 25 °C. As before, the concentration units are omitted in the K_w value.

$$K_w = [H_3O^+][OH^-]$$
$$= [1.0 \times 10^{-7}][1.0 \times 10^{-7}] = 1.0 \times 10^{-14}$$

ENGAGE

Why is the $[H_3O^+]$ equal to the $[OH^-]$ in pure water?

Neutral, Acidic, and Basic Solutions

The K_w value (1.0×10^{-14}) applies to any aqueous solution at 25 °C because all aqueous solutions contain both H_3O^+ and OH^- (see **FIGURE 14.3**). When the $[H_3O^+]$ and $[OH^-]$ in a solution are equal, the solution is **neutral**. However, most solutions are not neutral; they have different concentrations of $[H_3O^+]$ and $[OH^-]$. If acid is added to water, there is an increase in $[H_3O^+]$ and a decrease in $[OH^-]$, which makes an acidic solution. If base is added, $[OH^-]$ increases and $[H_3O^+]$ decreases, which gives a basic solution. However, for any aqueous solution, whether it is neutral, acidic, or basic, the product $[H_3O^+][OH^-]$ is equal to K_w (1.0×10^{-14}) at 25 °C (see **TABLE 14.6**).

ENGAGE

Why is the product of the concentrations of H_3O^+ and OH^- in a solution equal to 1.0×10^{-14}?

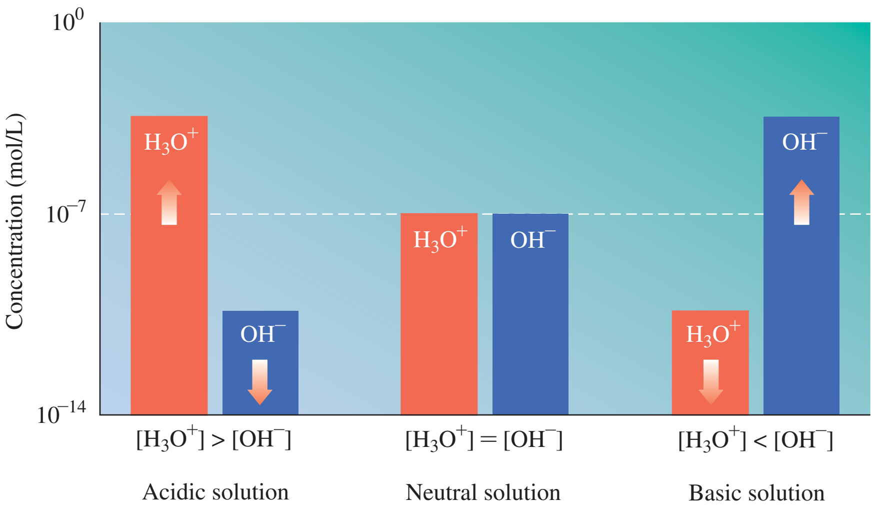


FIGURE 14.3 ▶ In a neutral solution, $[H_3O^+]$ and $[OH^-]$ are equal. In acidic solutions, the $[H_3O^+]$ is greater than the $[OH^-]$. In basic solutions, the $[OH^-]$ is greater than the $[H_3O^+]$.

Q Is a solution that has a $[H_3O^+]$ of 1.0×10^{-3} M acidic, basic, or neutral?

TABLE 14.6 Examples of $[H_3O^+]$ and $[OH^-]$ in Neutral, Acidic, and Basic Solutions

Type of Solution	$[H_3O^+]$	$[OH^-]$	K_w
Neutral	1.0×10^{-7} M	1.0×10^{-7} M	1.0×10^{-14}
Acidic	1.0×10^{-2} M	1.0×10^{-12} M	1.0×10^{-14}
Acidic	2.5×10^{-5} M	4.0×10^{-10} M	1.0×10^{-14}
Basic	1.0×10^{-8} M	1.0×10^{-6} M	1.0×10^{-14}
Basic	5.0×10^{-11} M	2.0×10^{-4} M	1.0×10^{-14}

Using the K_w to Calculate $[H_3O^+]$ and $[OH^-]$ in a Solution

If we know the $[H_3O^+]$ of a solution, we can use the K_w to calculate the $[OH^-]$. If we know the $[OH^-]$ of a solution, we can calculate $[H_3O^+]$ from their relationship in the K_w, as shown in Sample Problem 14.6.

$$K_w = [H_3O^+][OH^-]$$

$$[OH^-] = \frac{K_w}{[H_3O^+]} \qquad [H_3O^+] = \frac{K_w}{[OH^-]}$$

CORE CHEMISTRY SKILL
Calculating $[H_3O^+]$ and $[OH^-]$ in Solutions

ENGAGE
If you know the $[H_3O^+]$ of a solution, how do you use the K_w to calculate the $[OH^-]$?

SAMPLE PROBLEM 14.6 Calculating the $[H_3O^+]$ of a Solution

A vinegar solution has a $[OH^-] = 5.0 \times 10^{-12}$ M at 25 °C. What is the $[H_3O^+]$ of the vinegar solution? Is the solution acidic, basic, or neutral?

TRY IT FIRST

SOLUTION

STEP 1 State the given and needed quantities.

ANALYZE THE PROBLEM	Given	Need	Connect
	$[OH^-] = 5.0 \times 10^{-12}$ M	$[H_3O^+]$	$K_w = [H_3O^+][OH^-]$

STEP 2 Write the K_w for water and solve for the unknown $[H_3O^+]$.

$$K_w = [H_3O^+][OH^-] = 1.0 \times 10^{-14}$$

Solve for $[H_3O^+]$ by dividing both sides by $[OH^-]$.

$$\frac{K_w}{[OH^-]} = \frac{[H_3O^+]\cancel{[OH^-]}}{\cancel{[OH^-]}} = \frac{1.0 \times 10^{-14}}{[OH^-]}$$

$$[H_3O^+] = \frac{1.0 \times 10^{-14}}{[OH^-]}$$

STEP 3 Substitute the known $[OH^-]$ into the equation and calculate.

$$[H_3O^+] = \frac{1.0 \times 10^{-14}}{[5.0 \times 10^{-12}]} = 2.0 \times 10^{-3}\ \text{M}$$

Because the $[H_3O^+]$ of 2.0×10^{-3} M is larger than the $[OH^-]$ of 5.0×10^{-12} M, the solution is acidic.

STUDY CHECK 14.6

What is the $[H_3O^+]$ of an ammonia cleaning solution with $[OH^-] = 4.0 \times 10^{-4}$ M? Is the solution acidic, basic, or neutral?

ANSWER

$[H_3O^+] = 2.5 \times 10^{-11}$ M, basic

Guide to Calculating $[H_3O^+]$ and $[OH^-]$ in Aqueous Solutions

STEP 1 State the given and needed quantities.

STEP 2 Write the K_w for water and solve for the unknown $[H_3O^+]$ or $[OH^-]$.

STEP 3 Substitute the known $[H_3O^+]$ or $[OH^-]$ into the equation and calculate.

QUESTIONS AND PROBLEMS

14.5 Dissociation of Water

LEARNING GOAL Use the water dissociation expression to calculate the $[H_3O^+]$ and $[OH^-]$ in an aqueous solution.

14.33 Why are the concentrations of H_3O^+ and OH^- equal in pure water?

14.34 What is the meaning and value of K_w at 25 °C?

14.35 In an acidic solution, how does the concentration of H_3O^+ compare to the concentration of OH^-?

14.36 If a base is added to pure water, why does the $[H_3O^+]$ decrease?

14.37 Indicate whether each of the following solutions is acidic, basic, or neutral:
- **a.** $[H_3O^+] = 2.0 \times 10^{-5}$ M
- **b.** $[H_3O^+] = 1.4 \times 10^{-9}$ M
- **c.** $[OH^-] = 8.0 \times 10^{-3}$ M
- **d.** $[OH^-] = 3.5 \times 10^{-10}$ M

14.38 Indicate whether each of the following solutions is acidic, basic, or neutral:
- **a.** $[H_3O^+] = 6.0 \times 10^{-12}$ M
- **b.** $[H_3O^+] = 1.4 \times 10^{-4}$ M
- **c.** $[OH^-] = 5.0 \times 10^{-12}$ M
- **d.** $[OH^-] = 4.5 \times 10^{-2}$ M

14.39 Calculate the $[H_3O^+]$ of each aqueous solution with the following $[OH^-]$:
- **a.** coffee, 1.0×10^{-9} M
- **b.** soap, 1.0×10^{-6} M
- **c.** cleanser, 2.0×10^{-5} M
- **d.** lemon juice, 4.0×10^{-13} M

14.40 Calculate the $[H_3O^+]$ of each aqueous solution with the following $[OH^-]$:
- **a.** NaOH, 1.0×10^{-2} M
- **b.** milk of magnesia, 1.0×10^{-5} M
- **c.** aspirin, 1.8×10^{-11} M
- **d.** seawater, 2.5×10^{-6} M

Applications

14.41 Calculate the $[OH^-]$ of each aqueous solution with the following $[H_3O^+]$:
- **a.** stomach acid, 4.0×10^{-2} M
- **b.** urine, 5.0×10^{-6} M
- **c.** orange juice, 2.0×10^{-4} M
- **d.** bile, 7.9×10^{-9} M

14.42 Calculate the $[OH^-]$ of each aqueous solution with the following $[H_3O^+]$:
- **a.** baking soda, 1.0×10^{-8} M
- **b.** blood, 4.2×10^{-8} M
- **c.** milk, 5.0×10^{-7} M
- **d.** pancreatic juice, 4.0×10^{-9} M

14.6 The pH Scale

LEARNING GOAL Calculate pH from $[H_3O^+]$; given the pH, calculate the $[H_3O^+]$ and $[OH^-]$ of a solution.

In the environment, the acidity, or pH, of rain, water, and soil can have significant effects. When rain becomes too acidic, it can dissolve marble statues and accelerate the corrosion of metals. In lakes and ponds, the acidity of water can affect the ability of plants and fish to survive. The acidity of soil around plants affects their growth. If the soil pH is too acidic or too basic, the roots of the plant cannot take up some nutrients. Most plants thrive in soil with a nearly neutral pH, although certain plants, such as orchids, camellias, and blueberries, require a more acidic soil.

Personnel working in food processing, medicine, agriculture, spa and pool maintenance, soap manufacturing, and wine making measure the $[H_3O^+]$ and $[OH^-]$ of solutions. Although we have expressed H_3O^+ and OH^- as molar concentrations, it is more convenient to describe the acidity of solutions using the *pH scale*. On this scale, a number between 0 and 14 represents the H_3O^+ concentration for common solutions. A neutral solution has a pH of 7.0 at 25 °C. An acidic solution has a pH less than 7.0; a basic solution has a pH greater than 7.0 (see **FIGURE 14.4**).

If soil is too acidic, nutrients are not absorbed by crops. Then lime ($CaCO_3$), which acts as a base, may be added to increase the soil pH.

Acidic solution	pH < 7.0	$[H_3O^+] > 1.0 \times 10^{-7}$ M
Neutral solution	pH = 7.0	$[H_3O^+] = 1.0 \times 10^{-7}$ M
Basic solution	pH > 7.0	$[H_3O^+] < 1.0 \times 10^{-7}$ M

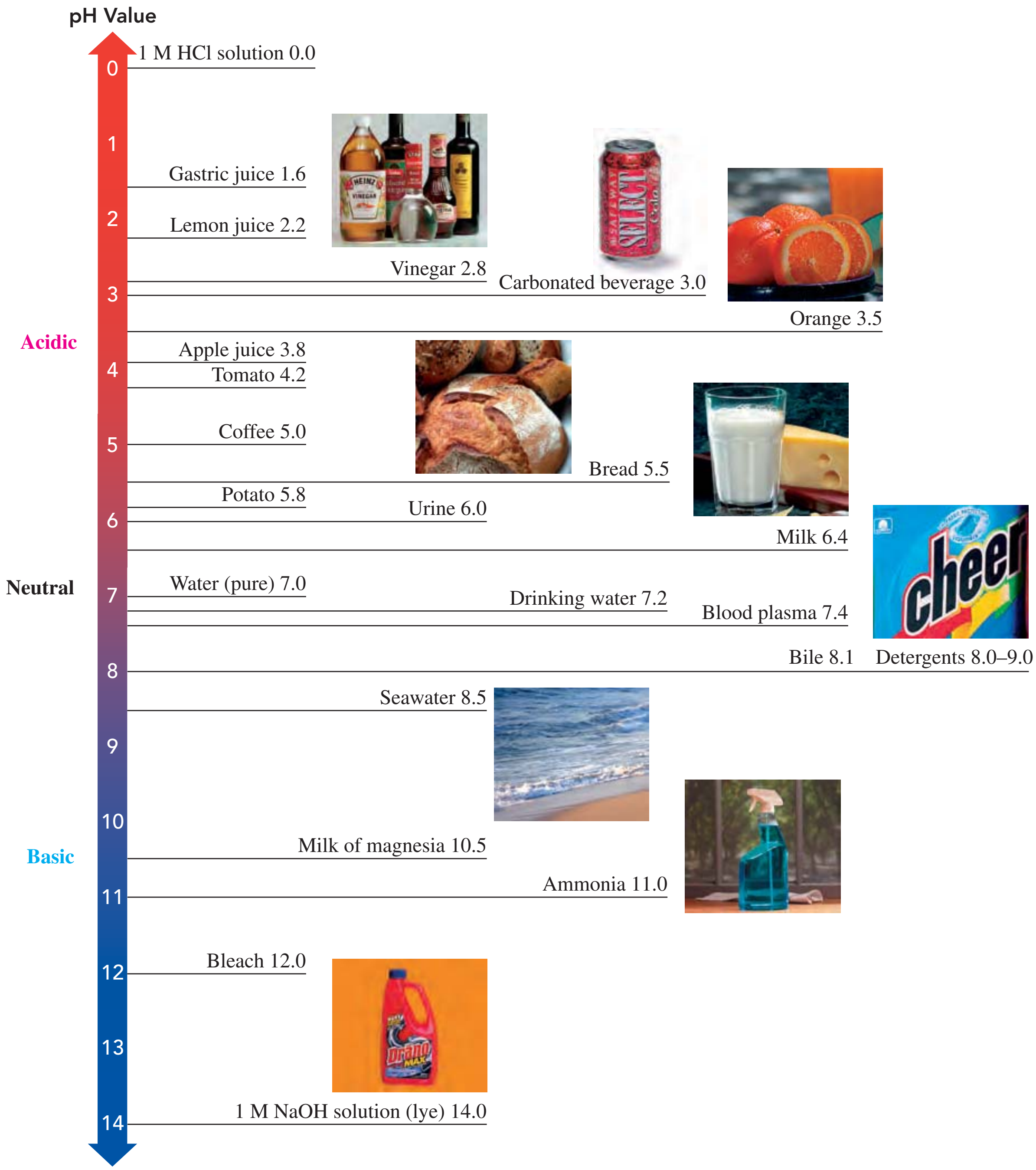


FIGURE 14.4 ▶ On the pH scale, values below 7.0 are acidic, a value of 7.0 is neutral, and values above 7.0 are basic.

Q Is apple juice an acidic, a basic, or a neutral solution?

When we relate acidity and pH, we are using an inverse relationship, which is when one component increases while the other component decreases. When an acid is added to pure water, the $[H_3O^+]$ (acidity) of the solution increases but its pH decreases. When a base is added to pure water, it becomes more basic, which means its acidity decreases and the pH increases.

In the laboratory, a pH meter is commonly used to determine the pH of a solution. There are also various indicators and pH papers that turn specific colors when placed in solutions of different pH values. The pH is found by comparing the color on the test paper or the color of the solution to a color chart (see **FIGURE 14.5**).

FIGURE 14.5 ▶ The pH of a solution can be determined using **(a)** a pH meter, **(b)** pH paper, and **(c)** indicators that turn different colors corresponding to different pH values.

If a pH meter reads 4.00, is the solution acidic, basic, or neutral?

(a) (b) (c)

A dipstick is used to measure the pH of a urine sample.

SAMPLE PROBLEM 14.7 pH of Solutions

Consider the pH of the following body fluids:

Body Fluid	pH
Stomach acid	1.4
Pancreatic juice	8.4
Sweat	4.8
Urine	5.3
Cerebrospinal fluid	7.3

a. Place the pH values of the body fluids on the list in order of most acidic to most basic.
b. Which body fluid has the highest $[H_3O^+]$?

TRY IT FIRST

SOLUTION

a. The most acidic body fluid is the one with the lowest pH, and the most basic is the body fluid with the highest pH: stomach acid (1.4), sweat (4.8), urine (5.3), cerebrospinal fluid (7.3), pancreatic juice (8.4).
b. The body fluid with the highest $[H_3O^+]$ would have the lowest pH value, which is stomach acid.

STUDY CHECK 14.7

Which body fluid has the highest $[OH^-]$?

ANSWER

The body fluid with the highest $[OH^-]$ would have the highest pH value, which is pancreatic juice.

KEY MATH SKILL
Calculating pH from $[H_3O^+]$

Calculating the pH of Solutions

The pH scale is a logarithmic scale that corresponds to the $[H_3O^+]$ of aqueous solutions. Mathematically, **pH** is the negative logarithm (base 10) of the $[H_3O^+]$.

$$\text{pH} = -\log[H_3O^+]$$

Essentially, the negative powers of 10 in the molar concentrations are converted to positive numbers. For example, a lemon juice solution with $[H_3O^+] = 1.0 \times 10^{-2}$ M has a pH of 2.00. This can be calculated using the pH equation:

$$\text{pH} = -\log[1.0 \times 10^{-2}]$$
$$\text{pH} = -(-2.00)$$
$$= 2.00$$

The number of *decimal places* in the pH value is the same as the number of significant figures in the $[H_3O^+]$. The number to the left of the decimal point in the pH value is the power of 10.

$$[H_3O^+] = \mathbf{1.0} \times 10^{-2} \qquad pH = 2.\mathbf{00}$$

Two SFs → Two SFs

Because pH is a log scale, a change of one pH unit corresponds to a tenfold change in $[H_3O^+]$. It is important to note that the pH decreases as the $[H_3O^+]$ increases. For example, a solution with a pH of 2.00 has a $[H_3O^+]$ that is ten times greater than a solution with a pH of 3.00 and 100 times greater than a solution with a pH of 4.00. The pH of a solution is calculated from the $[H_3O^+]$ by using the *log* key and changing the sign as shown in Sample Problem 14.8.

ENGAGE

Explain why 6.00 but not 6.0 is the correct pH for $[H_3O^+] = 1.0 \times 10^{-6}$ M.

SAMPLE PROBLEM 14.8 Calculating pH from $[H_3O^+]$

Aspirin, which is acetylsalicylic acid, was the first nonsteroidal anti-inflammatory drug (NSAID) used to alleviate pain and fever. If a solution of aspirin has a $[H_3O^+] = 1.7 \times 10^{-3}$ M, what is the pH of the solution?

Aspirin, acetylsalicylic acid, is a weak acid.

TRY IT FIRST

SOLUTION

STEP 1 State the given and needed quantities.

ANALYZE THE PROBLEM	Given	Need	Connect
	$[H_3O^+] = 1.7 \times 10^{-3}$ M	pH	pH equation

STEP 2 Enter the $[H_3O^+]$ into the pH equation and calculate.

$$pH = -\log[H_3O^+] = -\log[1.7 \times 10^{-3}]$$

Calculator Procedure | **Calculator Display**

1.7 [EE or EXP] [+/–] 3 [log] [+/–] [=] or [+/–] [log] 1.7 [EE or EXP] [+/–] 3 [=] | 2.769551079

Be sure to check the instructions for your calculator. Different calculators can have different methods for pH calculation.

Guide to Calculating pH of an Aqueous Solution

STEP 1 State the given and needed quantities.

STEP 2 Enter the $[H_3O^+]$ into the pH equation and calculate.

STEP 3 Adjust the number of SFs on the *right* of the decimal point.

STEP 3 Adjust the number of SFs on the *right* of the decimal point. In a pH value, the number to the *left* of the decimal point is an *exact* number derived from the power of 10. Thus, the two SFs in the coefficient determine that there are two SFs after the decimal point in the pH value.

Coefficient		Power of ten	
1.7 (Two SFs)	$\times$	10^{-3} M (Exact)	$pH = -\log[1.7 \times 10^{-3}] = 2.77$ (2: Exact; 77: Two SFs)

STUDY CHECK 14.8

What is the pH of bleach with $[H_3O^+] = 4.2 \times 10^{-12}$ M?

ANSWER

pH = 11.38

When we need to calculate the pH from $[OH^-]$, we use the K_w to calculate $[H_3O^+]$, place it in the pH equation, and calculate the pH of the solution as shown in Sample Problem 14.9.

SAMPLE PROBLEM 14.9 Calculating pH from $[OH^-]$

What is the pH of an ammonia solution with $[OH^-] = 3.7 \times 10^{-3}$ M?

TRY IT FIRST

SOLUTION

STEP 1 State the given and needed quantities.

ANALYZE THE PROBLEM	Given	Need	Connect
	$[OH^-] = 3.7 \times 10^{-3}$ M	pH	$K_w = [H_3O^+][OH^-]$, pH equation

STEP 2 Enter the $[H_3O^+]$ into the pH equation and calculate. Because $[OH^-]$ is given for the ammonia solution, we have to calculate $[H_3O^+]$. Using the water dissociation expression, K_w, we divide both sides by $[OH^-]$ to obtain $[H_3O^+]$.

$$K_w = [H_3O^+][OH^-] = 1.0 \times 10^{-14}$$

$$\frac{K_w}{[OH^-]} = \frac{[H_3O^+]\cancel{[OH^-]}}{\cancel{[OH^-]}}$$

$$[H_3O^+] = \frac{1.0 \times 10^{-14}}{[3.7 \times 10^{-3}]} = 2.7 \times 10^{-12}\text{ M}$$

Now, we enter the $[H_3O^+]$ into the pH equation.

$$pH = -\log[H_3O^+] = -\log[2.7 \times 10^{-12}]$$

Calculator Procedure	Calculator Display
2.7 [EE or EXP] [+/–] 12 [log] [+/–] [=]	11.56863624
or	
[+/–] [log] 2.7 [EE or EXP] [+/–] 12 [=]	

STEP 3 Adjust the number of SFs on the *right* of the decimal point.

$\mathbf{2.7} \times 10^{-12}$ M — Two SFs

pH = 11.**57** — Two SFs to the *right* of the decimal point

STUDY CHECK 14.9

Calculate the pH of a sample of bile that has $[OH^-] = 1.3 \times 10^{-6}$ M.

ANSWER

pH = 8.11

pOH

The **pOH** scale is similar to the pH scale except that pOH is associated with the $[OH^-]$ of an aqueous solution.

$$\text{pOH} = -\log[OH^-]$$

Solutions with high $[OH^-]$ have low pOH values; solutions with low $[OH^-]$ have high pOH values. In any aqueous solution, the sum of the pH and pOH is equal to 14.00, which is the negative logarithm of the K_w.

$$\text{pH} + \text{pOH} = 14.00$$

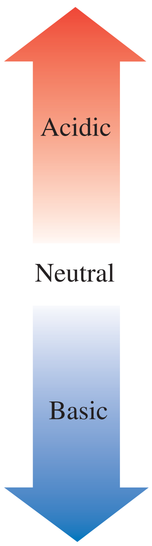
For example, if the pH of a solution is 3.50, the pOH can be calculated as follows:

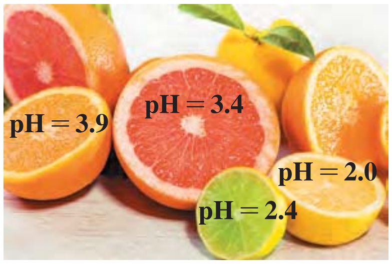
$$\text{pH} + \text{pOH} = 14.00$$
$$\text{pOH} = 14.00 - \text{pH} = 14.00 - 3.50 = 10.50$$

A comparison of $[H_3O^+]$, $[OH^-]$, and their corresponding pH and pOH values is given in **TABLE 14.7**.

TABLE 14.7 A Comparison of pH and pOH Values at 25 °C, $[H_3O^+]$, and $[OH^-]$

pH	$[H_3O^+]$	$[OH^-]$	pOH
0	10^{0}	10^{-14}	14
1	10^{-1}	10^{-13}	13
2	10^{-2}	10^{-12}	12
3	10^{-3}	10^{-11}	11
4	10^{-4}	10^{-10}	10
5	10^{-5}	10^{-9}	9
6	10^{-6}	10^{-8}	8
7	10^{-7}	10^{-7}	7
8	10^{-8}	10^{-6}	6
9	10^{-9}	10^{-5}	5
10	10^{-10}	10^{-4}	4
11	10^{-11}	10^{-3}	3
12	10^{-12}	10^{-2}	2
13	10^{-13}	10^{-1}	1
14	10^{-14}	10^{0}	0





Acids produce the sour taste of the fruits we eat.

KEY MATH SKILL
Calculating $[H_3O^+]$ from pH

Calculating $[H_3O^+]$ from pH

If we are given the pH of the solution and asked to determine the $[H_3O^+]$, we need to reverse the calculation of pH.

$$[H_3O^+] = 10^{-pH}$$

For example, if the pH of a solution is 3.0, we can substitute it into this equation. The number of significant figures in $[H_3O^+]$ is equal to the number of decimal places in the pH value.

$$[H_3O^+] = 10^{-pH} = 10^{-3.0} = 1 \times 10^{-3}\ M$$

For pH values that are not whole numbers, the calculation requires the use of the *10^x* key, which is usually a *2nd function* key. On some calculators, this operation is done using the inverse log equation as shown in Sample Problem 14.10.

SAMPLE PROBLEM 14.10 Calculating $[H_3O^+]$ from pH

Calculate $[H_3O^+]$ for a urine sample, which has a pH of 7.5.

TRY IT FIRST

SOLUTION

Guide to Calculating $[H_3O^+]$ from pH

STEP 1 State the given and needed quantities.

STEP 2 Enter the pH value into the inverse log equation and calculate.

STEP 3 Adjust the SFs for the coefficient.

STEP 1 State the given and needed quantities.

ANALYZE THE PROBLEM	Given	Need	Connect
	pH = 7.5	$[H_3O^+]$	$[H_3O^+] = 10^{-pH}$

STEP 2 Enter the pH value into the inverse log equation and calculate.

$$[H_3O^+] = 10^{-pH} = 10^{-7.5}$$

Calculator Procedure	Calculator Display
2nd log +/− 7.5 = or 7.5 +/− 2nd log =	3.16227766E−08

Be sure to check the instructions for your calculator. Different calculators can have different methods for this calculation.

STEP 3 **Adjust the SFs for the coefficient.** Because the pH value 7.5 has one digit to the *right* of the decimal point, the coefficient for $[H_3O^+]$ is written with one SF.

$$[H_3O^+] = \underbrace{3}_{\text{One SF}} \times 10^{-8}\ M$$

STUDY CHECK

What are the $[H_3O^+]$ and $[OH^-]$ of Diet Coke that has a pH of 3.17?

ANSWER

$[H_3O^+] = 6.8 \times 10^{-4}$ M, $[OH^-] = 1.5 \times 10^{-11}$ M

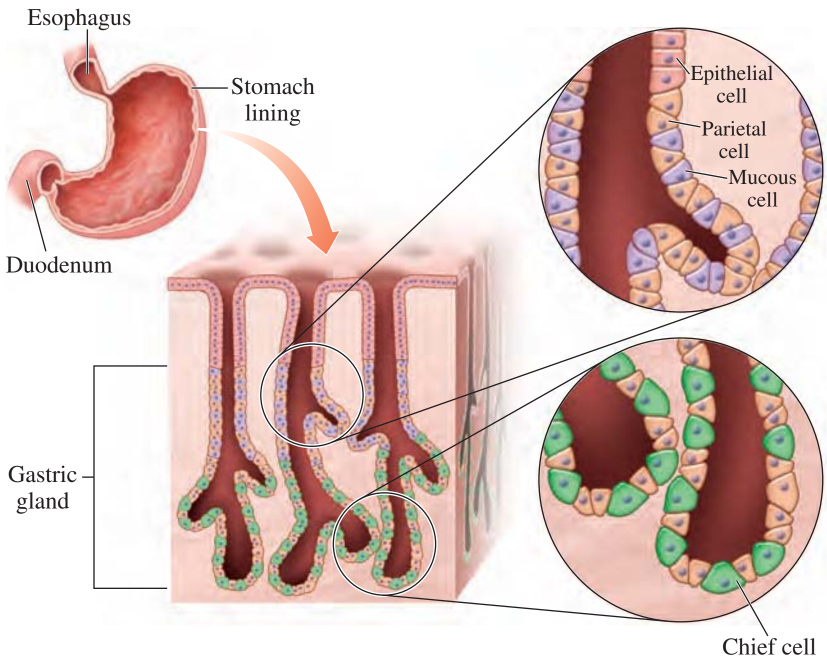
The pH of Diet Coke is 3.17.

CHEMISTRY LINK TO HEALTH

Stomach Acid, HCl

Gastric acid, which contains HCl, is produced by parietal cells that line the stomach. When the stomach expands with the intake of food, the gastric glands begin to secrete a strongly acidic solution of HCl. In a single day, a person may secrete 2000 mL of gastric juice, which contains hydrochloric acid, mucins, and the enzymes pepsin and lipase.

The HCl in the gastric juice activates a digestive enzyme from the chief cells called *pepsinogen* to form *pepsin*, which breaks down proteins in food entering the stomach. The secretion of HCl continues until the stomach has a pH of about 2, which is the optimum for activating the digestive enzymes without ulcerating the stomach lining. In addition, the low pH destroys bacteria that reach the stomach. Normally, large quantities of viscous mucus are secreted within the stomach to protect its lining from acid and enzyme damage. Gastric acid may also form under conditions of stress when the nervous system activates the production of HCl. As the contents of the stomach move into the small intestine, cells produce bicarbonate that neutralizes the gastric acid until the pH is about 5.



Parietal cells in the lining of the stomach secrete gastric acid HCl.

QUESTIONS AND PROBLEMS

14.6 The pH Scale

LEARNING GOAL Calculate pH from $[H_3O^+]$; given the pH, calculate the $[H_3O^+]$ and $[OH^-]$ of a solution.

14.43 Why does a neutral solution have a pH of 7.0?

14.44 If you know the $[OH^-]$, how can you determine the pH of a solution?

14.45 State whether each of the following solutions is acidic, basic, or neutral:
a. blood plasma, pH 7.38 **b.** vinegar, pH 2.8
c. drain cleaner, pOH 2.8 **d.** coffee, pH 5.52
e. tomatoes, pH 4.2 **f.** chocolate cake, pH 7.6

14.46 State whether each of the following solutions is acidic, basic, or neutral:
a. soda, pH 3.22 **b.** shampoo, pOH 8.3
c. laundry detergent, pOH 4.56 **d.** rain, pH 5.8
e. honey, pH 3.9 **f.** cheese, pH 4.9

14.47 A solution with a pH of 3 is 10 times more acidic than a solution with pH 4. Explain.

14.48 A solution with a pH of 10 is 100 times more basic than a solution with pH 8. Explain.

14.49 Calculate the pH of each solution given the following:
a. $[H_3O^+] = 1 \times 10^{-4}$ M **b.** $[H_3O^+] = 3 \times 10^{-9}$ M
c. $[OH^-] = 1 \times 10^{-5}$ M **d.** $[OH^-] = 2.5 \times 10^{-11}$ M
e. $[H_3O^+] = 6.7 \times 10^{-8}$ M **f.** $[OH^-] = 8.2 \times 10^{-4}$ M

14.50 Calculate the pOH of each solution given the following:
a. $[H_3O^+] = 1 \times 10^{-8}$ M **b.** $[H_3O^+] = 5 \times 10^{-6}$ M
c. $[OH^-] = 1 \times 10^{-2}$ M **d.** $[OH^-] = 8.0 \times 10^{-3}$ M
e. $[H_3O^+] = 4.7 \times 10^{-2}$ M **f.** $[OH^-] = 3.9 \times 10^{-6}$ M

Applications

14.51 Complete the following table:

$[H_3O^+]$	$[OH^-]$	pH	pOH	Acidic, Basic, or Neutral?
	1.0×10^{-6} M			
		3.49		
2.8×10^{-5} M				
			2.00	

14.52 Complete the following table:

$[H_3O^+]$	$[OH^-]$	pH	pOH	Acidic, Basic, or Neutral?
		10.00		
				Neutral
			5.66	
6.4×10^{-12} M				

14.53 A patient with severe metabolic acidosis has a blood plasma pH of 6.92. What is the $[H_3O^+]$ of the blood plasma?

14.54 A patient with respiratory alkalosis has a blood plasma pH of 7.58. What is the $[H_3O^+]$ of the blood plasma?

14.7 Reactions of Acids and Bases

LEARNING GOAL Write balanced equations for reactions of acids with metals, carbonates or bicarbonates, and bases.

Typical reactions of acids and bases include the reactions of acids with metals, carbonates or bicarbonates, and bases. For example, when you drop an antacid tablet in water, the bicarbonate ion and citric acid in the tablet react to produce carbon dioxide bubbles, water, and salt. A **salt** is an ionic compound that does not have H^+ as the cation or OH^- as the anion.

Magnesium reacts rapidly with acid and forms H_2 gas and a salt of magnesium.

Acids and Metals

Acids react with certain metals to produce hydrogen gas (H_2) and a salt. Active metals include potassium, sodium, calcium, magnesium, aluminum, zinc, iron, and tin. In these single replacement reactions, the metal ion replaces the hydrogen in the acid.

$$\underbrace{Mg(s)}_{\text{Metal}} + \underbrace{2HCl(aq)}_{\text{Acid}} \longrightarrow \underbrace{H_2(g)}_{\text{Hydrogen}} + \underbrace{MgCl_2(aq)}_{\text{Salt}}$$

$$\underbrace{Zn(s)}_{\text{Metal}} + \underbrace{2HNO_3(aq)}_{\text{Acid}} \longrightarrow \underbrace{H_2(g)}_{\text{Hydrogen}} + \underbrace{Zn(NO_3)_2(aq)}_{\text{Salt}}$$

Acids React with Carbonates or Bicarbonates

When an acid is added to a carbonate or bicarbonate, the products are carbon dioxide gas, water, and a salt. The acid reacts with CO_3^{2-} or HCO_3^- to produce carbonic acid (H_2CO_3), which breaks down rapidly to CO_2 and H_2O.

$$\underbrace{2HCl(aq)}_{\text{Acid}} + \underbrace{Na_2CO_3(aq)}_{\text{Carbonate}} \longrightarrow \underbrace{CO_2(g)}_{\text{Carbon dioxide}} + \underbrace{H_2O(l)}_{\text{Water}} + \underbrace{2NaCl(aq)}_{\text{Salt}}$$

$$\underbrace{HBr(aq)}_{\text{Acid}} + \underbrace{NaHCO_3(aq)}_{\text{Bicarbonate}} \longrightarrow \underbrace{CO_2(g)}_{\text{Carbon dioxide}} + \underbrace{H_2O(l)}_{\text{Water}} + \underbrace{NaBr(aq)}_{\text{Salt}}$$

When sodium bicarbonate (baking soda) reacts with an acid (vinegar), the products are carbon dioxide gas, water, and a salt.

Acids and Hydroxides: Neutralization

Neutralization is a reaction between a strong or weak acid with a strong base to produce water and a salt. The H^+ of the acid and the OH^- of the base combine to form water. The salt is the combination of the cation from the base and the anion from the acid. We can write the following equation for the neutralization reaction between HCl and NaOH:

$$\underbrace{HCl(aq)}_{\text{Acid}} + \underbrace{NaOH(aq)}_{\text{Base}} \longrightarrow \underbrace{H_2O(l)}_{\text{Water}} + \underbrace{NaCl(aq)}_{\text{Salt}}$$

If we write the strong acid HCl and the strong base NaOH as ions, we see that H^+ combines with OH^- to form water, leaving the ions Na^+ and Cl^- in solution.

$$H^+(aq) + Cl^-(aq) + Na^+(aq) + OH^-(aq) \longrightarrow H_2O(l) + Na^+(aq) + Cl^-(aq)$$

When we omit the ions that do not change during the reaction (*spectator ions*), we obtain the *net ionic equation.*

$$H^+(aq) + \cancel{Cl^-(aq)} + \cancel{Na^+(aq)} + OH^-(aq) \longrightarrow H_2O(l) + \cancel{Na^+(aq)} + \cancel{Cl^-(aq)}$$

CORE CHEMISTRY SKILL Writing Equations for Reactions of Acids and Bases

The net ionic equation for the neutralization of H^+ and OH^- to form H_2O is

$$H^+(aq) + OH^-(aq) \longrightarrow H_2O(l) \quad \text{Net ionic equation}$$

Balancing Neutralization Equations

In a neutralization reaction, one H^+ always reacts with one OH^-. Therefore, a neutralization equation may need coefficients to balance the H^+ from the acid with the OH^- from the base as shown in Sample Problem 14.11.

SAMPLE PROBLEM 14.11 Balancing Equations for Acids

Write the balanced equation for the neutralization of $HCl(aq)$ and $Ba(OH)_2(s)$.

TRY IT FIRST

SOLUTION

STEP 1 Write the reactants and products.

$$HCl(aq) + Ba(OH)_2(s) \longrightarrow H_2O(l) + \text{salt}$$

STEP 2 Balance the H^+ in the acid with the OH^- in the base. Placing a coefficient of 2 in front of the HCl provides $2H^+$ for the $2OH^-$ from $Ba(OH)_2$.

$$2HCl(aq) + Ba(OH)_2(s) \longrightarrow H_2O(l) + \text{salt}$$

STEP 3 Balance the H_2O with the H^+ and the OH^-. Use a coefficient of 2 in front of H_2O to balance $2H^+$ and $2OH^-$.

$$2HCl(aq) + Ba(OH)_2(s) \longrightarrow 2H_2O(l) + \text{salt}$$

STEP 4 Write the formula of the salt from the remaining ions. Use the ions Ba^{2+} and $2Cl^-$ and write the formula for the salt as $BaCl_2$.

$$2HCl(aq) + Ba(OH)_2(s) \longrightarrow 2H_2O(l) + BaCl_2(aq)$$

Guide to Balancing an Equation for Neutralization

STEP 1 Write the reactants and products.

STEP 2 Balance the H^+ in the acid with the OH^- in the base.

STEP 3 Balance the H_2O with the H^+ and the OH^-.

STEP 4 Write the formula of the salt from the remaining ions.

STUDY CHECK 14.11

Write the balanced equation for the neutralization of $H_2SO_4(aq)$ and $LiOH(aq)$.

ANSWER

$$H_2SO_4(aq) + 2LiOH(aq) \longrightarrow 2H_2O(l) + Li_2SO_4(aq)$$

CHEMISTRY LINK TO HEALTH

Antacids

Antacids are substances used to neutralize excess stomach acid (HCl). Some antacids are mixtures of aluminum hydroxide and magnesium hydroxide. These hydroxides are not very soluble in water, so the levels of available OH^- are not damaging to the intestinal tract. However, aluminum hydroxide has the side effects of producing constipation and binding phosphate in the intestinal tract, which may cause weakness and loss of appetite. Magnesium hydroxide has a laxative effect. These side effects are less likely when a combination of the antacids is used.

$$Al(OH)_3(s) + 3HCl(aq) \longrightarrow 3H_2O(l) + AlCl_3(aq)$$

$$Mg(OH)_2(s) + 2HCl(aq) \longrightarrow 2H_2O(l) + MgCl_2(aq)$$

Antacids neutralize excess stomach acid.

Some antacids use calcium carbonate to neutralize excess stomach acid. About 10% of the calcium is absorbed into the bloodstream, where it elevates the level of serum calcium. Calcium carbonate is not recommended for patients who have peptic ulcers or a tendency to form kidney stones, which typically consist of an insoluble calcium salt.

$$CaCO_3(s) + 2HCl(aq) \longrightarrow CO_2(g) + H_2O(l) + CaCl_2(aq)$$

Still other antacids contain sodium bicarbonate. This type of antacid neutralizes excess gastric acid, increases blood pH, but also elevates sodium levels in the body fluids. It also is not recommended in the treatment of peptic ulcers.

$$NaHCO_3(s) + HCl(aq) \longrightarrow CO_2(g) + H_2O(l) + NaCl(aq)$$

The neutralizing substances in some antacid preparations are given in **TABLE 14.8**.

TABLE 14.8 Basic Compounds in Some Antacids

Antacid	Base(s)
Amphojel	$Al(OH)_3$
Milk of magnesia	$Mg(OH)_2$
Mylanta, Maalox, Di-Gel, Gelusil, Riopan	$Mg(OH)_2$, $Al(OH)_3$
Bisodol, Rolaids	$CaCO_3$, $Mg(OH)_2$
Titralac, Tums, Pepto-Bismol	$CaCO_3$
Alka-Seltzer	$NaHCO_3$, $KHCO_3$

QUESTIONS AND PROBLEMS

14.7 Reactions of Acids and Bases

LEARNING GOAL Write balanced equations for reactions of acids with metals, carbonates or bicarbonates, and bases.

14.55 Complete and balance the equation for each of the following reactions:
- **a.** $ZnCO_3(s) + HBr(aq) \longrightarrow$
- **b.** $Zn(s) + HCl(aq) \longrightarrow$
- **c.** $HCl(aq) + NaHCO_3(s) \longrightarrow$
- **d.** $H_2SO_4(aq) + Mg(OH)_2(s) \longrightarrow$

14.56 Complete and balance the equation for each of the following reactions:
- **a.** $KHCO_3(s) + HBr(aq) \longrightarrow$
- **b.** $Ca(s) + H_2SO_4(aq) \longrightarrow$
- **c.** $H_2SO_4(aq) + Ca(OH)_2(s) \longrightarrow$
- **d.** $Na_2CO_3(s) + H_2SO_4(aq) \longrightarrow$

14.57 Balance each of the following neutralization reactions:
- **a.** $HCl(aq) + Mg(OH)_2(s) \longrightarrow H_2O(l) + MgCl_2(aq)$
- **b.** $H_3PO_4(aq) + LiOH(aq) \longrightarrow H_2O(l) + Li_3PO_4(aq)$

14.58 Balance each of the following neutralization reactions:
- **a.** $HNO_3(aq) + Ba(OH)_2(s) \longrightarrow H_2O(l) + Ba(NO_3)_2(aq)$
- **b.** $H_2SO_4(aq) + Al(OH)_3(s) \longrightarrow H_2O(l) + Al_2(SO_4)_3(aq)$

14.59 Write a balanced equation for the neutralization of each of the following:
- **a.** $H_2SO_4(aq)$ and $NaOH(aq)$
- **b.** $HCl(aq)$ and $Fe(OH)_3(s)$
- **c.** $H_2CO_3(aq)$ and $Mg(OH)_2(s)$

14.60 Write a balanced equation for the neutralization of each of the following:
- **a.** $H_3PO_4(aq)$ and $NaOH(aq)$
- **b.** $HI(aq)$ and $LiOH(aq)$
- **c.** $HNO_3(aq)$ and $Ca(OH)_2(s)$

14.8 Acid–Base Titration

LEARNING GOAL Calculate the molarity or volume of an acid or base solution from titration information.

Suppose we need to find the molarity of a solution of HCl, which has an unknown concentration. We can do this by a laboratory procedure called **titration** in which we neutralize an acid sample with a known amount of base. In a titration, we place a measured volume of the acid in a flask and add a few drops of an **indicator**, such as phenolphthalein. An indicator is a compound that dramatically changes color when pH of the solution changes. In an acidic solution, phenolphthalein is colorless. Then we fill a buret with a NaOH solution of known molarity and carefully add NaOH solution to neutralize the acid in the flask (see **FIGURE 14.6**). We know that neutralization has taken place when the phenolphthalein in the solution changes from colorless to pink. This is called the neutralization **endpoint**. From the measured volume of the NaOH solution and its molarity, we calculate the number of moles of NaOH, the moles of acid, and use the measured volume of acid to calculate its concentration.

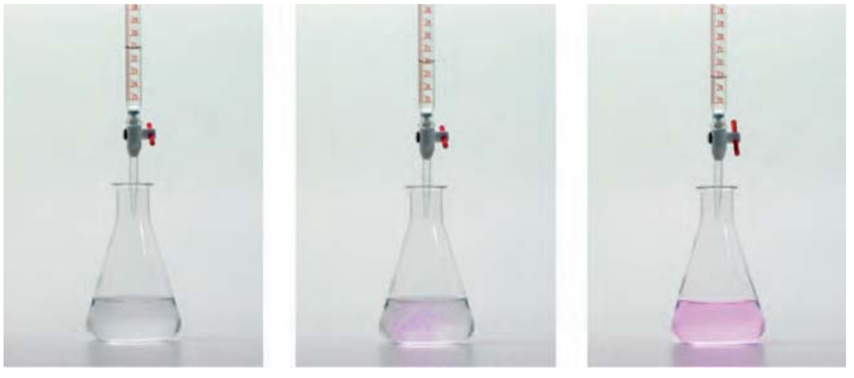


FIGURE 14.6 ▶ The titration of an acid. A known volume of an acid is placed in a flask with an indicator and titrated with a measured volume of a base solution, such as NaOH, to the neutralization endpoint.

Q What data is needed to determine the molarity of the acid in the flask?

SAMPLE PROBLEM 14.12 Titration of an Acid

If 16.3 mL of a 0.185 M $Sr(OH)_2$ solution is used to titrate the HCl in 0.0250 L of gastric juice, what is the molarity of the HCl solution?

$$Sr(OH)_2(aq) + 2HCl(aq) \longrightarrow 2H_2O(l) + SrCl_2(aq)$$

TRY IT FIRST

SOLUTION

STEP 1 State the given and needed quantities and concentrations.

ANALYZE THE PROBLEM	Given	Need	Connect
	0.0250 L of HCl solution, 16.3 mL of 0.185 M $Sr(OH)_2$ solution	molarity of the HCl solution	molarity, mole–mole factor
	Neutralization Equation		
	$Sr(OH)_2(aq) + 2HCl(aq) \longrightarrow 2H_2O(l) + SrCl_2(aq)$		

STEP 2 Write a plan to calculate the molarity.

mL of $Sr(OH)_2$ solution → (Metric factor) → L of $Sr(OH)_2$ solution → (Molarity) → moles of $Sr(OH)_2$ → (Mole–mole factor) → moles of HCl → (Divide by liters) → molarity of HCl solution

STEP 3 State equalities and conversion factors, including concentrations.

$$1 \text{ L of } Sr(OH)_2 \text{ solution} = 1000 \text{ mL of } Sr(OH)_2 \text{ solution}$$
$$\frac{1000 \text{ mL } Sr(OH)_2 \text{ solution}}{1 \text{ L } Sr(OH)_2 \text{ solution}} \text{ and } \frac{1 \text{ L } Sr(OH)_2 \text{ solution}}{1000 \text{ mL } Sr(OH)_2 \text{ solution}}$$

$$1 \text{ L of } Sr(OH)_2 \text{ solution} = 0.185 \text{ mol of } Sr(OH)_2$$
$$\frac{0.185 \text{ mol } Sr(OH)_2}{1 \text{ L } Sr(OH)_2 \text{ solution}} \text{ and } \frac{1 \text{ L } Sr(OH)_2 \text{ solution}}{0.185 \text{ mol } Sr(OH)_2}$$

$$2 \text{ mol of HCl} = 1 \text{ mol of } Sr(OH)_2$$
$$\frac{2 \text{ mol HCl}}{1 \text{ mol } Sr(OH)_2} \text{ and } \frac{1 \text{ mol } Sr(OH)_2}{2 \text{ mol HCl}}$$

STEP 4 Set up the problem to calculate the needed quantity.

$$16.3 \cancel{\text{ mL } Sr(OH)_2 \text{ solution}} \times \frac{1 \cancel{\text{ L } Sr(OH)_2 \text{ solution}}}{1000 \cancel{\text{ mL } Sr(OH)_2 \text{ solution}}} \times \frac{0.185 \cancel{\text{ mol } Sr(OH)_2}}{1 \cancel{\text{ L } Sr(OH)_2 \text{ solution}}} \times \frac{2 \text{ mol HCl}}{1 \cancel{\text{ mol } Sr(OH)_2}}$$

$$= 0.006\,03 \text{ mol of HCl}$$

$$\text{molarity of HCl solution} = \frac{0.006\,03 \text{ mol HCl}}{0.0250 \text{ L HCl solution}} = 0.241 \text{ M HCl solution}$$

STUDY CHECK 14.12

What is the molarity of an HCl solution if 28.6 mL of a 0.175 M NaOH solution is needed to titrate a 25.0-mL sample of the HCl solution?

ANSWER

0.200 M HCl solution

CORE CHEMISTRY SKILL
Calculating Molarity or Volume of an Acid or Base in a Titration

Guide to Calculations for an Acid–Base Titration

STEP 1 State the given and needed quantities and concentrations.

STEP 2 Write a plan to calculate the molarity or volume.

STEP 3 State equalities and conversion factors, including concentrations.

STEP 4 Set up the problem to calculate the needed quantity.

Interactive Video

Acid–Base Titration

QUESTIONS AND PROBLEMS

14.8 Acid–Base Titration

LEARNING GOAL Calculate the molarity or volume of an acid or base solution from titration information.

14.61 If you need to determine the molarity of a formic acid solution, $HCHO_2$, how would you proceed?

$$HCHO_2(aq) + H_2O(l) \rightleftarrows H_3O^+(aq) + CHO_2^-(aq)$$

14.62 If you need to determine the molarity of an acetic acid solution, $HC_2H_3O_2$, how would you proceed?

$$HC_2H_3O_2(aq) + H_2O(l) \rightleftarrows H_3O^+(aq) + C_2H_3O_2^-(aq)$$

14.63 What is the molarity of a solution of HCl if 5.00 mL of the HCl solution is titrated with 28.6 mL of a 0.145 M NaOH solution?

$$HCl(aq) + NaOH(aq) \longrightarrow H_2O(l) + NaCl(aq)$$

14.64 What is the molarity of an acetic acid solution if 25.0 mL of the $HC_2H_3O_2$ solution is titrated with 29.7 mL of a 0.205 M KOH solution?

$$HC_2H_3O_2(aq) + KOH(aq) \longrightarrow H_2O(l) + KC_2H_3O_2(aq)$$

14.65 If 38.2 mL of a 0.163 M KOH solution is required to neutralize completely 25.0 mL of a solution of H_2SO_4, what is the molarity of the H_2SO_4 solution?

$$H_2SO_4(aq) + 2KOH(aq) \longrightarrow 2H_2O(l) + K_2SO_4(aq)$$

14.66 A solution of 0.162 M NaOH is used to titrate 25.0 mL of a solution of H_2SO_4. If 32.8 mL of the NaOH solution is required to reach the endpoint, what is the molarity of the H_2SO_4 solution?

$$H_2SO_4(aq) + 2NaOH(aq) \longrightarrow 2H_2O(l) + Na_2SO_4(aq)$$

14.67 A solution of 0.204 M NaOH is used to titrate 50.0 mL of a 0.0224 M H_3PO_4 solution. What volume, in milliliters, of the NaOH solution is required?

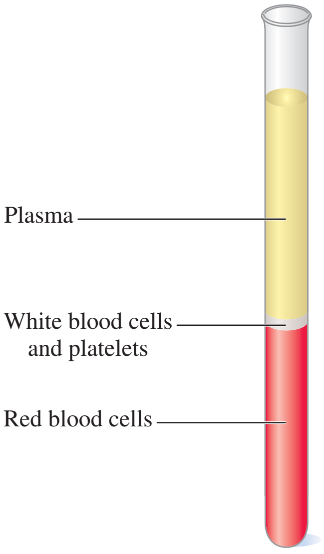
$$H_3PO_4(aq) + 3NaOH(aq) \longrightarrow 3H_2O(l) + Na_3PO_4(aq)$$

14.68 A solution of 0.312 M KOH is used to titrate 15.0 mL of a 0.186 M H_3PO_4 solution. What volume, in milliliters, of the KOH solution is required?

$$H_3PO_4(aq) + 3KOH(aq) \longrightarrow 3H_2O(l) + K_3PO_4(aq)$$

14.9 Buffers

LEARNING GOAL Describe the role of buffers in maintaining the pH of a solution; calculate the pH of a buffer.



Whole blood consists of plasma, white blood cells and platelets, and red blood cells.

ENGAGE

Why does a buffer require the presence of a weak acid or weak base and the salt of that weak acid or weak base?

The lungs and the kidneys are the primary organs that regulate the pH of body fluids, including blood and urine. Major changes in the pH of the body fluids can severely affect biological activities within the cells. *Buffers* are present to prevent large fluctuations in pH.

The pH of water and most solutions changes drastically when a small amount of acid or base is added. However, when an acid or a base is added to a buffer solution, there is little change in pH. A **buffer solution** maintains the pH of a solution by neutralizing small amounts of added acid or base. In the human body, whole blood contains plasma, white blood cells and platelets, and red blood cells. Blood plasma contains buffers that maintain a consistent pH of about 7.4. If the pH of the blood plasma goes slightly above or below 7.4, changes in our oxygen levels and our metabolic processes can be drastic enough to cause death. Even though we obtain acids and bases from foods and cellular reactions, the buffers in the body absorb those compounds so effectively that the pH of our blood plasma remains essentially unchanged (see **FIGURE 14.7**).

In a buffer, an acid must be present to react with any OH^- that is added, and a base must be available to react with any added H_3O^+. However, that acid and base must not neutralize each other. Therefore, a combination of an acid–base conjugate pair is used in buffers. Most buffer solutions consist of nearly equal concentrations of a weak acid and a salt containing its conjugate base. Buffers may also contain a weak base and the salt of the weak base, which contains its conjugate acid.

For example, a typical buffer can be made from the weak acid acetic acid ($HC_2H_3O_2$) and its salt, sodium acetate ($NaC_2H_3O_2$). As a weak acid, acetic acid dissociates slightly in water to form H_3O^+ and a very small amount of $C_2H_3O_2^-$. The addition of its salt, sodium acetate, provides a much larger concentration of acetate ion ($C_2H_3O_2^-$), which is necessary for its buffering capability.

$$\underset{\text{Large amount}}{HC_2H_3O_2(aq)} + H_2O(l) \rightleftarrows H_3O^+(aq) + \underset{\text{Large amount}}{C_2H_3O_2^-(aq)}$$

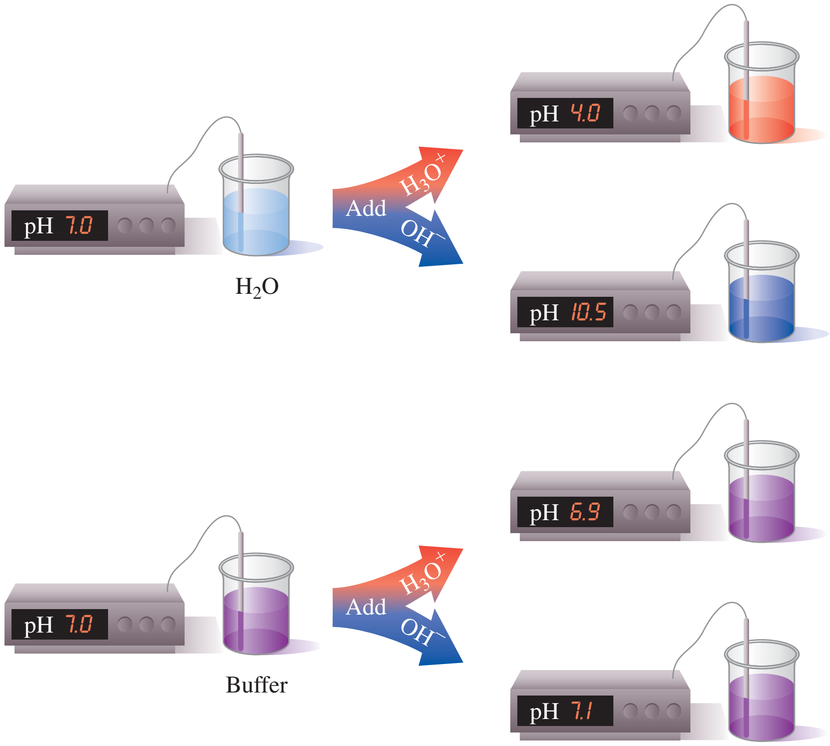


FIGURE 14.7 ▶ Adding an acid or a base to water changes the pH drastically, but a buffer resists pH change when small amounts of acid or base are added.

Why does the pH change several pH units when acid is added to water, but not when acid is added to a buffer?

We can now describe how this buffer solution maintains the $[H_3O^+]$. When a small amount of acid is added, the additional H_3O^+ combines with the acetate ion, $C_2H_3O_2^-$, causing the equilibrium to shift in the direction of the reactants, acetic acid and water. There will be a slight decrease in the $[C_2H_3O_2^-]$ and a slight increase in the $[HC_2H_3O_2]$, but both the $[H_3O^+]$ and pH are maintained (see **FIGURE 14.8**).

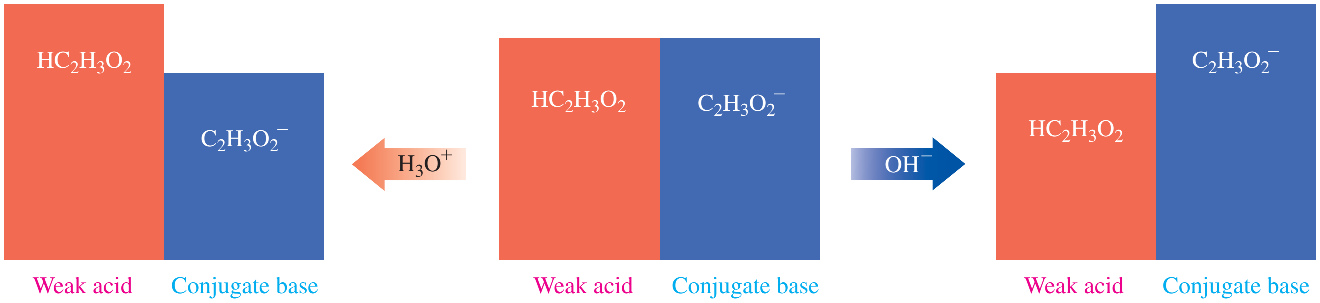


FIGURE 14.8 ▶ The buffer described here consists of about equal concentrations of acetic acid ($HC_2H_3O_2$) and its conjugate base acetate ion ($C_2H_3O_2^-$). Adding H_3O^+ to the buffer neutralizes some $C_2H_3O_2^-$, whereas adding OH^- neutralizes some $HC_2H_3O_2$. The pH of the solution is maintained as long as the added amount of acid or base is small compared to the concentrations of the buffer components.

How does this acetic acid–acetate ion buffer maintain pH?

$$HC_2H_3O_2(aq) + H_2O(l) \longleftarrow H_3O^+(aq) + C_2H_3O_2^-(aq)$$

Equilibrium shifts in the direction of the reactants

ENGAGE

Which part of a buffer neutralizes any H_3O^+ that is added?

If a small amount of base is added to this same buffer solution, it is neutralized by the acetic acid, $HC_2H_3O_2$, which shifts the equilibrium in the direction of the products acetate ion and water. The $[HC_2H_3O_2]$ decreases slightly and the $[C_2H_3O_2^-]$ increases slightly, but again the $[H_3O^+]$ and thus the pH of the solution are maintained.

$$HC_2H_3O_2(aq) + OH^-(aq) \longrightarrow H_2O(l) + C_2H_3O_2^-(aq)$$

Equilibrium shifts in the direction of the products

CORE CHEMISTRY SKILL
Calculating the pH of a Buffer

Calculating the pH of a Buffer

By rearranging the K_a expression to give $[H_3O^+]$, we can obtain the ratio of the acetic acid/acetate buffer.

$$K_a = \frac{[H_3O^+][C_2H_3O_2^-]}{[HC_2H_3O_2]}$$

Solving for $[H_3O^+]$ gives:

$$[H_3O^+] = K_a \times \frac{[HC_2H_3O_2] \longleftarrow \text{Weak acid}}{[C_2H_3O_2^-] \longleftarrow \text{Conjugate base}}$$

In this rearrangement of K_a, the weak acid is in the numerator and the conjugate base in the denominator. We can now calculate the $[H_3O^+]$ and pH for an acetic acid buffer as shown in Sample Problem 14.13.

SAMPLE PROBLEM 14.13 Calculating the pH of a Buffer

The K_a for acetic acid, $HC_2H_3O_2$, is 1.8×10^{-5}. What is the pH of a buffer prepared with 1.0 M $HC_2H_3O_2$ and 1.0 M $C_2H_3O_2^-$?

$$HC_2H_3O_2(aq) + H_2O(l) \rightleftarrows H_3O^+(aq) + C_2H_3O_2^-(aq)$$

TRY IT FIRST

Guide to Calculating pH of a Buffer

STEP 1 State the given and needed quantities.

STEP 2 Write the K_a expression and rearrange for $[H_3O^+]$.

STEP 3 Substitute [HA] and $[A^-]$ into the K_a expression.

STEP 4 Use $[H_3O^+]$ to calculate pH.

SOLUTION

STEP 1 State the given and needed quantities.

	Given	Need	Connect
ANALYZE THE PROBLEM	1.0 M $HC_2H_3O_2$, 1.0 M $C_2H_3O_2^-$	pH	K_a expression
	Equation		
	$HC_2H_3O_2(aq) + H_2O(l) \rightleftarrows H_3O^+(aq) + C_2H_3O_2^-(aq)$		

STEP 2 Write the K_a expression and rearrange for $[H_3O^+]$.

$$K_a = \frac{[H_3O^+][C_2H_3O_2^-]}{[HC_2H_3O_2]}$$

$$[H_3O^+] = K_a \times \frac{[HC_2H_3O_2]}{[C_2H_3O_2^-]}$$

STEP 3 Substitute [HA] and $[A^-]$ into the K_a expression.

$$[H_3O^+] = 1.8 \times 10^{-5} \times \frac{[1.0]}{[1.0]}$$

$$[H_3O^+] = 1.8 \times 10^{-5}\ \text{M}$$

STEP 4 Use [H_3O^+] to calculate pH. Placing the [H_3O^+] into the pH equation gives the pH of the buffer.

$$\text{pH} = -\log[1.8 \times 10^{-5}] = 4.74$$

STUDY CHECK 14.13

One of the conjugate acid–base pairs that buffers the blood plasma is $H_2PO_4^-/HPO_4^{2-}$, which has a K_a of 6.2×10^{-8}. What is the pH of a buffer that is prepared from 0.10 M $H_2PO_4^-$ and 0.50 M HPO_4^{2-}?

ANSWER

pH = 7.91

Because K_a is a constant at a given temperature, the [H_3O^+] is determined by the [$HC_2H_3O_2$]/[$C_2H_3O_2^-$] ratio. As long as the addition of small amounts of either acid or base changes the ratio of [$HC_2H_3O_2$]/[$C_2H_3O_2^-$] only slightly, the changes in [H_3O^+] will be small and the pH will be maintained. If a large amount of acid or base is added, the *buffering capacity* of the system may be exceeded. Buffers can be prepared from conjugate acid–base pairs such as $H_2PO_4^-/HPO_4^{2-}$, HPO_4^{2-}/PO_4^{3-}, HCO_3^-/CO_3^{2-}, or NH_4^+/NH_3. The pH of the buffer solution will depend on the conjugate acid–base pair chosen.

ENGAGE

How would a solution composed of HPO_4^{2-} and PO_4^{3-} act as a buffer?

Using a common phosphate buffer for biological specimens, we can look at the effect of using different ratios of [$H_2PO_4^-$]/[HPO_4^{2-}] on the [H_3O^+] and pH. The K_a of $H_2PO_4^-$ is 6.2×10^{-8}. The equation and the [H_3O^+] are written as follows:

$$H_2PO_4^-(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + HPO_4^{2-}(aq)$$

$$[H_3O^+] = K_a \times \frac{[H_2PO_4^-]}{[HPO_4^{2-}]}$$

K_a	$\frac{[H_2PO_4^-]}{[HPO_4^{2-}]}$	Ratio	[H_3O^+]	pH
6.2×10^{-8}	$\frac{1.0\ M}{0.10\ M}$	$\frac{10}{1}$	6.2×10^{-7}	6.21
6.2×10^{-8}	$\frac{1.0\ M}{1.0\ M}$	$\frac{1}{1}$	6.2×10^{-8}	7.21
6.2×10^{-8}	$\frac{0.10\ M}{1.0\ M}$	$\frac{1}{10}$	6.2×10^{-9}	8.21

To prepare a phosphate buffer with a pH close to the pH of a biological sample, 7.4, we would choose concentrations that are about equal, such as 1.0 M $H_2PO_4^-$ and 1.0 M HPO_4^{2-}.

CHEMISTRY LINK TO HEALTH

Buffers in the Blood Plasma

The arterial blood plasma has a normal pH of 7.35 to 7.45. If changes in H_3O^+ lower the pH below 6.8 or raise it above 8.0, cells cannot function properly and death may result. In our cells, CO_2 is continually produced as an end product of cellular metabolism. Some CO_2 is carried to the lungs for elimination, and the rest dissolves in body fluids such as plasma and saliva, forming carbonic acid, H_2CO_3. As a weak acid, carbonic acid dissociates to give bicarbonate, HCO_3^-, and H_3O^+. More of the anion HCO_3^- is supplied by the kidneys to give an important buffer system in the body fluid—the H_2CO_3/HCO_3^- buffer.

$$CO_2(g) + H_2O(l) \rightleftharpoons H_2CO_3(aq) \rightleftharpoons H_3O^+(aq) + HCO_3^-(aq)$$

Excess H_3O^+ entering the body fluids reacts with the HCO_3^-, and excess OH^- reacts with the carbonic acid.

$$H_2CO_3(aq) + H_2O(l) \longleftarrow H_3O^+(aq) + HCO_3(aq)$$

Equilibrium shifts in the direction of the reactants

$$H_2CO_3(aq) + OH^-(aq) \longrightarrow H_2O(l) + HCO_3^-(aq)$$

Equilibrium shifts in the direction of the products

For carbonic acid, we can write the equilibrium expression as

$$K_a = \frac{[H_3O^+][HCO_3^-]}{[H_2CO_3]}$$

To maintain the normal blood plasma pH (7.35 to 7.45), the ratio of $[H_2CO_3]/[HCO_3^-]$ needs to be about 1 to 10, which is obtained by the concentrations in the blood plasma of 0.0024 M H_2CO_3 and 0.024 M HCO_3^-.

$$[H_3O^+] = K_a \times \frac{[H_2CO_3]}{[HCO_3^-]}$$

$$[H_3O^+] = 4.3 \times 10^{-7} \times \frac{[0.0024]}{[0.024]}$$

$$= 4.3 \times 10^{-7} \times 0.10 = 4.3 \times 10^{-8}\ M$$

$$pH = -\log[4.3 \times 10^{-8}] = 7.37$$

In the body, the concentration of carbonic acid is closely associated with the partial pressure of CO_2, P_{CO_2}. **TABLE 14.9** lists the normal values for arterial blood. If the CO_2 level rises, increasing $[H_2CO_3]$, the equilibrium shifts to produce more H_3O^+, which lowers the pH. This condition is called *acidosis*. Difficulty with ventilation or gas diffusion can lead to respiratory acidosis, which can happen in emphysema or when an accident or depressive drugs affect the medulla of the brain.

A lowering of the CO_2 level leads to a high blood pH, a condition called *alkalosis*. Excitement, trauma, or a high temperature may cause a person to hyperventilate, which expels large amounts of CO_2. As the partial pressure of CO_2 in the blood falls below normal, the equilibrium shifts from H_2CO_3 to CO_2 and H_2O. This shift decreases the $[H_3O^+]$ and raises the pH. The kidneys also regulate H_3O^+ and HCO_3^-, but they do so more slowly than the adjustment made by the lungs during ventilation.

TABLE 14.10 lists some of the conditions that lead to changes in the blood pH and some possible treatments.

TABLE 14.9 Normal Values for Blood Buffer in Arterial Blood

P_{CO_2}	40 mmHg
H_2CO_3	2.4 mmol/L of plasma
HCO_3^-	24 mmol/L of plasma
pH	7.35 to 7.45

TABLE 14.10 Acidosis and Alkalosis: Symptoms, Causes, and Treatments

Respiratory Acidosis: $CO_2\uparrow$ pH$\downarrow$	
Symptoms:	Failure to ventilate, suppression of breathing, disorientation, weakness, coma
Causes:	Lung disease blocking gas diffusion (e.g., emphysema, pneumonia, bronchitis, asthma); depression of respiratory center by drugs, cardiopulmonary arrest, stroke, poliomyelitis, or nervous system disorders
Treatment:	Correction of disorder, infusion of bicarbonate
Metabolic Acidosis: $H^+\uparrow$ pH$\downarrow$	
Symptoms:	Increased ventilation, fatigue, confusion
Causes:	Renal disease, including hepatitis and cirrhosis; increased acid production in diabetes mellitus, hyperthyroidism, alcoholism, and starvation; loss of alkali in diarrhea; acid retention in renal failure
Treatment:	Sodium bicarbonate given orally, dialysis for renal failure, insulin treatment for diabetic ketosis
Respiratory Alkalosis: $CO_2\downarrow$ pH$\uparrow$	
Symptoms:	Increased rate and depth of breathing, numbness, light-headedness, tetany
Causes:	Hyperventilation because of anxiety, hysteria, fever, exercise; reaction to drugs such as salicylate, quinine, and antihistamines; conditions causing hypoxia (e.g., pneumonia, pulmonary edema, heart disease)
Treatment:	Elimination of anxiety-producing state, rebreathing into a paper bag
Metabolic Alkalosis: $H^+\downarrow$ pH$\uparrow$	
Symptoms:	Depressed breathing, apathy, confusion
Causes:	Vomiting, diseases of the adrenal glands, ingestion of excess alkali
Treatment:	Infusion of saline solution, treatment of underlying diseases

QUESTIONS AND PROBLEMS

14.9 Buffers

LEARNING GOAL Describe the role of buffers in maintaining the pH of a solution; calculate the pH of a buffer.

14.69 Which of the following represents a buffer system? Explain.
- **a.** NaOH and NaCl
- **b.** H_2CO_3 and $NaHCO_3$
- **c.** HF and KF
- **d.** KCl and NaCl

14.70 Which of the following represents a buffer system? Explain.
- **a.** $HClO_2$
- **b.** $NaNO_3$
- **c.** $HC_2H_3O_2$ and $NaC_2H_3O_2$
- **d.** HCl and NaOH

14.71 Consider the buffer system of hydrofluoric acid, HF, and its salt, NaF.

$$HF(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + F^-(aq)$$

- **a.** The purpose of this buffer system is to:
  - **1.** maintain [HF]
  - **2.** maintain $[F^-]$
  - **3.** maintain pH
- **b.** The salt of the weak acid is needed to:
  - **1.** provide the conjugate base
  - **2.** neutralize added H_3O^+
  - **3.** provide the conjugate acid
- **c.** If OH^- is added, it is neutralized by:
  - **1.** the salt
  - **2.** H_2O
  - **3.** H_3O^+
- **d.** When H_3O^+ is added, the equilibrium shifts in the direction of the:
  - **1.** reactants
  - **2.** products
  - **3.** does not change

14.72 Consider the buffer system of nitrous acid, HNO_2, and its salt, $NaNO_2$.

$$HNO_2(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + NO_2^-(aq)$$

- **a.** The purpose of this buffer system is to:
  - **1.** maintain $[HNO_2]$
  - **2.** maintain $[NO_2^-]$
  - **3.** maintain pH
- **b.** The weak acid is needed to:
  - **1.** provide the conjugate base
  - **2.** neutralize added OH^-
  - **3.** provide the conjugate acid
- **c.** If H_3O^+ is added, it is neutralized by:
  - **1.** the salt
  - **2.** H_2O
  - **3.** OH^-
- **d.** When OH^- is added, the equilibrium shifts in the direction of the:
  - **1.** reactants
  - **2.** products
  - **3.** does not change

14.73 Nitrous acid has a K_a of 4.5×10^{-4}. What is the pH of a buffer solution containing 0.10 M HNO_2 and 0.10 M NO_2^-?

14.74 Acetic acid has a K_a of 1.8×10^{-5}. What is the pH of a buffer solution containing 0.15 M $HC_2H_3O_2$ and 0.15 M $C_2H_3O_2^-$?

14.75 Using Table 14.4 for K_a values, compare the pH of a HF buffer that contains 0.10 M HF and 0.10 M NaF with another HF buffer that contains 0.060 M HF and 0.120 M NaF.

14.76 Using Table 14.4 for K_a values, compare the pH of a H_2CO_3 buffer that contains 0.10 M H_2CO_3 and 0.10 M $NaHCO_3$ with another H_2CO_3 buffer that contains 0.15 M H_2CO_3 and 0.050 M $NaHCO_3$.

Applications

14.77 Why would the pH of your blood plasma increase if you breathe fast?

14.78 Why would the pH of your blood plasma decrease if you hold your breath?

14.79 Someone with kidney failure excretes urine with large amounts of HCO_3^-. How would this loss of HCO_3^- affect the pH of the blood plasma?

14.80 Someone with severe diabetes obtains energy by the breakdown of fats, which produce large amounts of acidic substances. How would this affect the pH of the blood plasma?

Follow Up

ACID REFLUX DISEASE

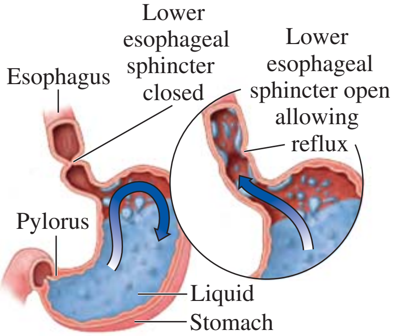
Larry has not been feeling well lately. He tells his doctor that he has discomfort and a burning feeling in his chest, and a sour taste in his throat and mouth. At times, Larry says he feels bloated after a big meal, has a dry cough, is hoarse, and sometimes has a sore throat. He has tried antacids, but they do not bring any relief.

The doctor tells Larry that he thinks he has acid reflux. At the top of the stomach there is a valve, the lower esophageal sphincter, that normally closes after food passes through it. However, if the valve does not close completely, acid produced in the stomach to digest food can move up into the esophagus, a condition called *acid reflux*. The acid, which is hydrochloric acid, HCl, is produced in the stomach to kill bacteria, microorganisms, and to activate the enzymes we need to break down food.

If acid reflux occurs, the strong acid HCl comes in contact with the lining of the esophagus, where it causes irritation and produces a burning feeling in the chest. Sometimes the pain in the chest is called *heartburn*. If the HCl reflux goes high enough to reach the throat, a sour taste may be noticed in the mouth. If Larry's symptoms occur three or more times a week, he may have a chronic condition known as *acid reflux disease* or *gastroesophageal reflux disease* (GERD).

Larry's doctor orders an *esophageal pH test* in which the amount of acid entering the esophagus from the stomach is measured over 24 h. A probe that measures the pH is inserted into the lower esophagus above the esophageal sphincter. The pH measurements indicate a reflux episode each time the pH drops to 4 or less.

In the 24-h period, Larry has several reflux episodes and his doctor determines that he has chronic GERD. He and Larry discuss treatment for GERD, which includes eating smaller meals, not lying down for 3 h after eating, making dietary changes, and losing weight. Antacids may be used to neutralize the acid coming up from the stomach. Other medications known as *proton pump inhibitors* (PPIs), such as Prilosec and Nexium, may be used to suppress the production of HCl in the stomach (gastric parietal cells), which raises the pH in the stomach to between 4 and 5, and gives the esophagus time to heal. Nexium may be given in oral doses of 40 mg once a day for 4 weeks. In severe GERD cases, an artificial valve may be created at the top of the stomach to strengthen the lower esophageal sphincter.



In acid reflux disease, the lower esophageal sphincter opens, allowing acidic fluid from the stomach to enter the esophagus.

Applications

14.81 At rest, the $[H_3O^+]$ of the stomach fluid is 2.0×10^{-4} M. What is the pH of the stomach fluid?

14.82 When food enters the stomach, HCl is released and the $[H_3O^+]$ of the stomach fluid rises to 4×10^{-2} M. What is the pH of the stomach fluid while eating?

14.83 In Larry's esophageal pH test, a pH value of 3.60 was recorded in the esophagus. What is the $[H_3O^+]$ in his esophagus?

14.84 After Larry had taken Nexium for 4 weeks, the pH in his stomach was raised to 4.52. What is the $[H_3O^+]$ in his stomach?

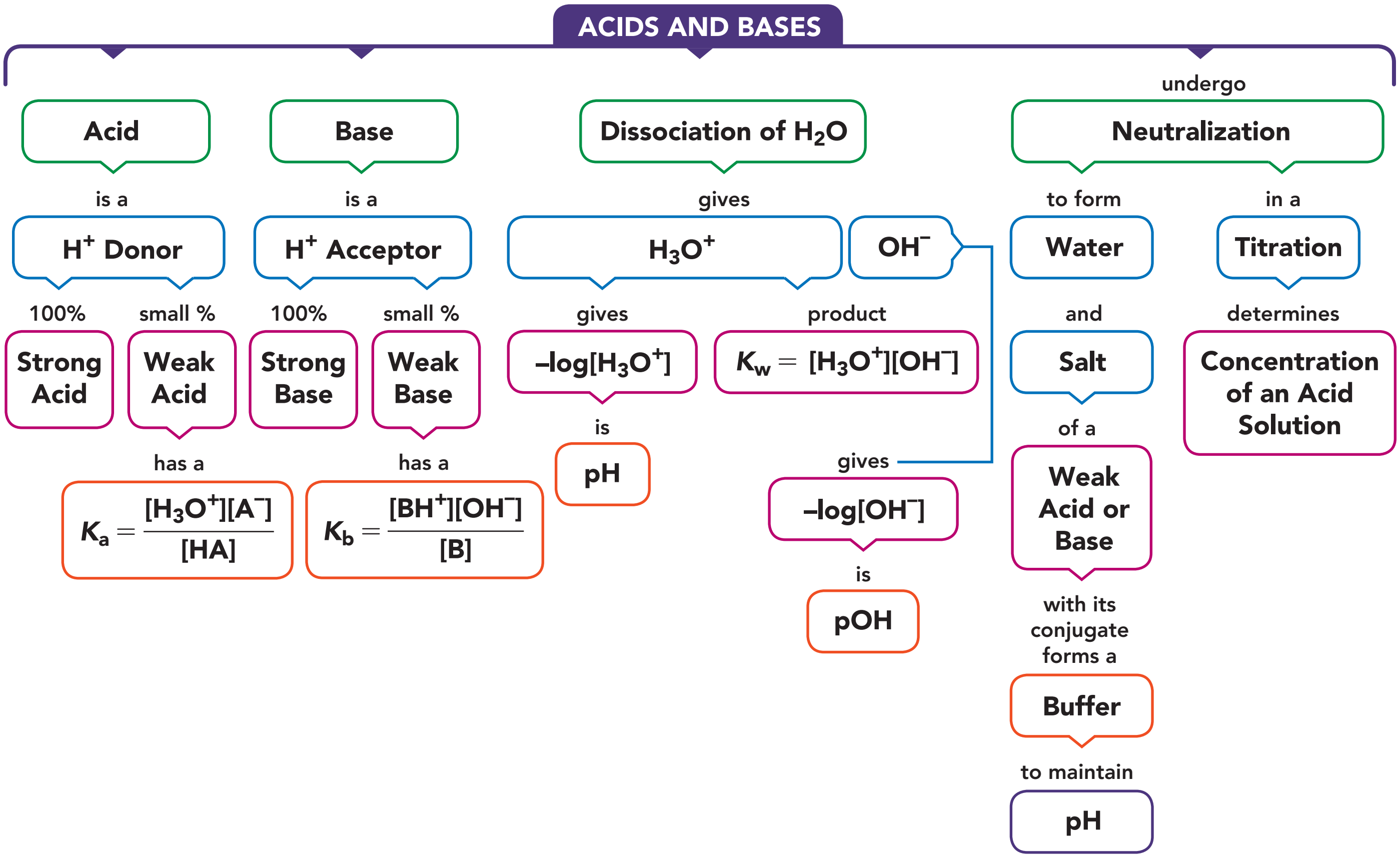
14.85 Write the balanced chemical equation for the neutralization reaction of stomach acid HCl with $CaCO_3$, an ingredient in some antacids.

14.86 Write the balanced chemical equation for the neutralization reaction of stomach acid HCl with $Al(OH)_3$, an ingredient in some antacids.

14.87 How many grams of $CaCO_3$ are required to neutralize 100. mL of stomach acid HCl, which is equivalent to 0.0400 M HCl?

14.88 How many grams of $Al(OH)_3$ are required to neutralize 150. mL of stomach acid HCl with a pH of 1.5?

CONCEPT MAP

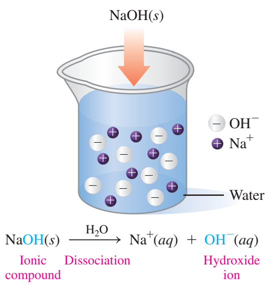


CHAPTER REVIEW

14.1 Acids and Bases

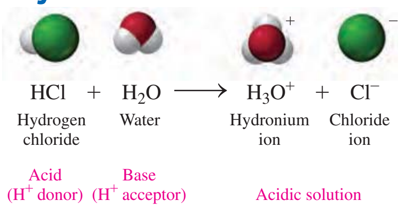
LEARNING GOAL Describe and name acids and bases.

- An Arrhenius acid produces H^+ and an Arrhenius base produces OH^- in aqueous solutions.
- Acids taste sour, may sting, and neutralize bases.
- Bases taste bitter, feel slippery, and neutralize acids.
- Acids containing a simple anion use a *hydro* prefix, whereas acids with oxygen-containing polyatomic anions are named as *ic* or *ous acids*.



14.2 Brønsted–Lowry Acids and Bases

LEARNING GOAL Identify conjugate acid–base pairs for Brønsted–Lowry acids and bases.



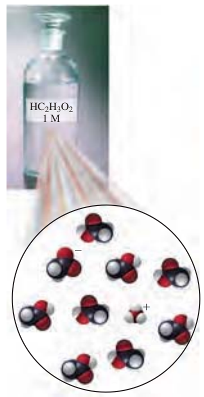
- According to the Brønsted–Lowry theory, acids are H^+ donors and bases are H^+ acceptors.
- A conjugate acid–base pair is related by the loss or gain of one H^+.
- For example, when the acid HF donates H^+, the F^- is its conjugate base. The other acid–base pair would be H_3O^+/H_2O.

$$HF(aq) + H_2O(l) \rightleftarrows H_3O^+(aq) + F^-(aq)$$

14.3 Strengths of Acids and Bases

LEARNING GOAL Write equations for the dissociation of strong and weak acids; identify the direction of reaction.

- Strong acids dissociate completely in water, and the H^+ is accepted by H_2O acting as a base.
- A weak acid dissociates slightly in water, producing only a small percentage of H_3O^+.
- Strong bases are hydroxides of Groups 1A (1) and 2A (2) that dissociate completely in water.
- An important weak base is ammonia, NH_3.



14.4 Dissociation Constants for Acids and Bases

LEARNING GOAL Write the dissociation expression for a weak acid or weak base.


$HCHO_2$


CHO_2^-

- In water, weak acids and weak bases produce only a few ions when equilibrium is reached.
- Weak acids have small K_a values whereas strong acids, which are essentially 100% dissociated, have very large K_a values.
- The reaction for a weak acid can be written as $HA + H_2O \rightleftarrows H_3O^+ + A^-$. The acid dissociation expression is written as

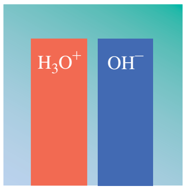
$$K_a = \frac{[H_3O^+][A^-]}{[HA]}.$$

- For a weak base, $B + H_2O \rightleftarrows BH^+ + OH^-$, the base dissociation expression is written as

$$K_b = \frac{[BH^+][OH^-]}{[B]}.$$

14.5 Dissociation of Water

LEARNING GOAL Use the water dissociation expression to calculate the $[H_3O^+]$ and $[OH^-]$ in an aqueous solution.



$[H_3O^+] = [OH^-]$
Neutral solution

- In pure water, a few water molecules transfer H^+ to other water molecules, producing small, but equal, amounts of $[H_3O^+]$ and $[OH^-]$.
- In pure water, the molar concentrations of H_3O^+ and OH^- are each 1.0×10^{-7} mol/L.
- The *water dissociation expression*, K_w, $[H_3O^+][OH^-] = 1.0 \times 10^{-14}$ at 25 °C.
- In acidic solutions, the $[H_3O^+]$ is greater than the $[OH^-]$.
- In basic solutions, the $[OH^-]$ is greater than the $[H_3O^+]$.

14.6 The pH Scale

LEARNING GOAL Calculate pH from $[H_3O^+]$; given the pH, calculate the $[H_3O^+]$ and $[OH^-]$ of a solution.

- The pH scale is a range of numbers typically from 0 to 14, which represents the $[H_3O^+]$ of the solution.
- A neutral solution has a pH of 7.0. In acidic solutions, the pH is below 7.0; in basic solutions, the pH is above 7.0.
- Mathematically, pH is the negative logarithm of the hydronium ion concentration,
  $pH = -\log[H_3O^+]$.
- The pOH is the negative log of the hydroxide ion concentration,
  $pOH = -\log[OH^-]$.
- The sum of the pH + pOH is 14.00.

14.7 Reactions of Acids and Bases

LEARNING GOAL Write balanced equations for reactions of acids with metals, carbonates, or bicarbonates, and bases.

- An acid reacts with a metal to produce hydrogen gas and a salt.
- The reaction of an acid with a carbonate or bicarbonate produces carbon dioxide, water, and a salt.
- In neutralization, an acid reacts with a base to produce water and a salt.

14.8 Acid–Base Titration

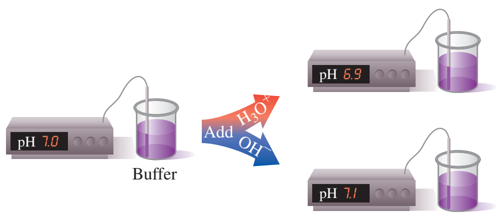
LEARNING GOAL Calculate the molarity or volume of an acid or base solution from titration information.

- In a titration, an acid sample is neutralized with a known amount of a base.
- From the volume and molarity of the base, the concentration of the acid is calculated.

14.9 Buffers

LEARNING GOAL Describe the role of buffers in maintaining the pH of a solution; calculate the pH of a buffer.

- A buffer solution resists changes in pH when small amounts of an acid or a base are added.
- A buffer contains either a weak acid and its salt or a weak base and its salt.
- In a buffer, the weak acid reacts with added OH^-, and the anion of the salt reacts with added H_3O^+.
- Most buffer solutions consist of nearly equal concentrations of a weak acid and a salt containing its conjugate base.
- The pH of a buffer is calculated by solving the K_a expression for $[H_3O^+]$.



KEY TERMS

acid A substance that dissolves in water and produces hydrogen ions (H^+), according to the Arrhenius theory. All acids are hydrogen ion donors, according to the Brønsted–Lowry theory.

acid dissociation expression, K_a The product of the ions from the dissociation of a weak acid divided by the concentration of the weak acid.

amphoteric Substances that can act as either an acid or a base in water.

base A substance that dissolves in water and produces hydroxide ions (OH^-) according to the Arrhenius theory. All bases are hydrogen ion acceptors, according to the Brønsted–Lowry theory.

base dissociation expression, K_b The product of the ions from the dissociation of a weak base divided by the concentration of the weak base.

Brønsted–Lowry acids and bases An acid is a hydrogen ion donor; a base is a hydrogen ion acceptor.

buffer solution A solution of a weak acid and its conjugate base or a weak base and its conjugate acid that maintains the pH by neutralizing added acid or base.

conjugate acid–base pair An acid and a base that differ by one H^+. When an acid donates a hydrogen ion, the product is its conjugate base, which is capable of accepting a hydrogen ion in the reverse reaction.

dissociation The separation of an acid or a base into ions in water.

endpoint The point at which an indicator changes color. For the indicator phenolphthalein, the color change occurs when the number of moles of OH^- is equal to the number of moles of H_3O^+ in the sample.

hydronium ion, H_3O^+ The ion formed by the attraction of a hydrogen ion, H^+, to a water molecule.

indicator A substance added to a titration sample that changes color when the pH of the solution changes.

neutral The term that describes a solution with equal concentrations of $[H_3O^+]$ and $[OH^-]$.

neutralization A reaction between an acid and a base to form water and a salt.

pH A measure of the $[H_3O^+]$ in a solution; $pH = -\log[H_3O^+]$.

pOH A measure of the $[OH^-]$ in a solution; $pOH = -\log[OH^-]$.

salt An ionic compound that contains a metal ion or NH_4^+ and a nonmetal or polyatomic ion other than OH^-.

strong acid An acid that completely dissociates in water.

strong base A base that completely dissociates in water.

titration The addition of base to an acid sample to determine the concentration of the acid.

water dissociation expression, K_w The product of $[H_3O^+]$ and $[OH^-]$ in solution; $K_w = [H_3O^+][OH^-]$.

weak acid An acid that is a poor donor of H^+ and dissociates only slightly in water.

weak base A base that is a poor acceptor of H^+ and produces only a small number of ions in water.

KEY MATH SKILLS

The chapter section containing each Key Math Skill is shown in parentheses at the end of each heading.

Calculating pH from $[H_3O^+]$ (14.6)

- The pH of a solution is calculated from the negative log of the $[H_3O^+]$.

$$pH = -\log[H_3O^+]$$

Example: What is the pH of a solution that has $[H_3O^+] = 2.4 \times 10^{-11}$ M?

Answer: We substitute the given $[H_3O^+]$ into the pH equation and calculate the pH.

$$\begin{aligned} pH &= -\log[H_3O^+] \\ &= -\log[2.4 \times 10^{-11}\ \text{M}] \\ &= 10.62 \end{aligned}$$

Two decimal places equal the two SFs in the $[H_3O^+]$ coefficient.

Calculating $[H_3O^+]$ from pH (14.6)

- The calculation of $[H_3O^+]$ from the pH is done by reversing the pH calculation using the negative pH.

$$[H_3O^+] = 10^{-pH}$$

Example: What is the $[H_3O^+]$ of a solution with a pH of 4.80?

Answer:

$$\begin{aligned} [H_3O^+] &= 10^{-pH} \\ &= 10^{-4.80} \\ &= 1.6 \times 10^{-5}\ \text{M} \end{aligned}$$

Two SFs in the $[H_3O^+]$ equal the two decimal places in the pH.

CORE CHEMISTRY SKILLS

The chapter section containing each Core Chemistry Skill is shown in parentheses at the end of each heading.

Identifying Conjugate Acid–Base Pairs (14.2)

- According to the Brønsted–Lowry theory, a conjugate acid–base pair consists of molecules or ions related by the loss of one H^+ by an acid, and the gain of one H^+ by a base.
- Every acid–base reaction contains two conjugate acid–base pairs because an H^+ is transferred in both the forward and reverse directions.
- When an acid such as HF loses one H^+, the conjugate base F^- is formed. When H_2O acts as a base, it gains one H^+, which forms its conjugate acid, H_3O^+.

Example: Identify the conjugate acid–base pairs in the following reaction:

$$H_2SO_4(aq) + H_2O(l) \longrightarrow HSO_4^-(aq) + H_3O^+(aq)$$

Answer:

$$\underset{\text{Acid}}{H_2SO_4(aq)} + \underset{\text{Base}}{H_2O(l)} \longrightarrow \underset{\text{Conjugate base}}{HSO_4^-(aq)} + \underset{\text{Conjugate acid}}{H_3O^+(aq)}$$

Conjugate acid–base pairs: H_2SO_4/HSO_4^- and H_3O^+/H_2O

Calculating $[H_3O^+]$ and $[OH^-]$ in Solutions (14.5)

- For all aqueous solutions, the product of $[H_3O^+]$ and $[OH^-]$ is equal to the water dissociation expression, K_w.

$$K_w = [H_3O^+][OH^-]$$

- Because pure water contains equal numbers of OH^- ions and H_3O^+ ions each with molar concentrations of 1.0×10^{-7} M, the numerical value of K_w is 1.0×10^{-14} at 25 °C.

$$\begin{aligned} K_w = [H_3O^+][OH^-] &= [1.0 \times 10^{-7}][1.0 \times 10^{-7}] \\ &= 1.0 \times 10^{-14} \end{aligned}$$

- If we know the $[H_3O^+]$ of a solution, we can use the K_w expression to calculate the $[OH^-]$. If we know the $[OH^-]$ of a solution, we can calculate the $[H_3O^+]$ using the K_w expression.

$$[OH^-] = \frac{K_w}{[H_3O^+]} \qquad [H_3O^+] = \frac{K_w}{[OH^-]}$$

Example: What is the $[OH^-]$ in a solution that has $[H_3O^+] = 2.4 \times 10^{-11}$ M? Is the solution acidic or basic?

Answer: We solve the K_w expression for $[OH^-]$ and substitute in the known values of K_w and $[H_3O^+]$.

$$[OH^-] = \frac{K_w}{[H_3O^+]} = \frac{1.0 \times 10^{-14}}{[2.4 \times 10^{-11}]} = 4.2 \times 10^{-4}\,\text{M}$$

Because the $[OH^-]$ is greater than the $[H_3O^+]$, this is a basic solution.

Writing Equations for Reactions of Acids and Bases (14.7)

- Acids react with certain metals to produce hydrogen gas (H_2) and a salt.

$$\underset{\text{Metal}}{Mg(s)} + \underset{\text{Acid}}{2HCl(aq)} \longrightarrow \underset{\text{Hydrogen}}{H_2(g)} + \underset{\text{Salt}}{MgCl_2(aq)}$$

- When an acid is added to a carbonate or bicarbonate, the products are carbon dioxide gas, water, and a salt.

$$\underset{\text{Acid}}{2HCl(aq)} + \underset{\text{Carbonate}}{Na_2CO_3(aq)} \longrightarrow \underset{\text{Carbon dioxide}}{CO_2(g)} + \underset{\text{Water}}{H_2O(l)} + \underset{\text{Salt}}{2NaCl(aq)}$$

- Neutralization is a reaction between a strong or weak acid and a strong base to produce water and a salt.

$$\underset{\text{Acid}}{HCl(aq)} + \underset{\text{Base}}{NaOH(aq)} \longrightarrow \underset{\text{Water}}{H_2O(l)} + \underset{\text{Salt}}{NaCl(aq)}$$

Example: Write the balanced chemical equation for the reaction of $ZnCO_3(s)$ and hydrobromic acid $HBr(aq)$.

Answer:

$$ZnCO_3(s) + 2HBr(aq) \longrightarrow CO_2(g) + H_2O(l) + ZnBr_2(aq)$$

Calculating Molarity or Volume of an Acid or Base in a Titration (14.8)

- In a titration, a measured volume of acid is neutralized by a NaOH solution of known molarity.
- From the measured volume of the NaOH solution required for titration and its molarity, the number of moles of NaOH, the moles of acid, and the concentration of the acid are calculated.

Example: A 15.0-mL sample of a H_2SO_4 solution is titrated with 24.0 mL of a 0.245 M NaOH solution. What is the molarity of the H_2SO_4 solution?

$$H_2SO_4(aq) + 2NaOH(aq) \longrightarrow 2H_2O(l) + Na_2SO_4(aq)$$

Answer:

$$24.0\ \cancel{\text{mL NaOH solution}} \times \frac{1\ \cancel{\text{L NaOH solution}}}{1000\ \cancel{\text{mL NaOH solution}}}$$

$$\times \frac{0.245\ \cancel{\text{mol NaOH}}}{1\ \cancel{\text{L NaOH solution}}} \times \frac{1\text{ mol }H_2SO_4}{2\ \cancel{\text{mol NaOH}}} = 0.002\,94\text{ mol of }H_2SO_4$$

$$\text{Molarity (M)} = \frac{0.002\,94\text{ mol }H_2SO_4}{0.0150\text{ L }H_2SO_4\text{ solution}} = 0.196\text{ M }H_2SO_4\text{ solution}$$

Calculating the pH of a Buffer (14.9)

- A buffer solution maintains pH by neutralizing small amounts of added acid or base.
- Most buffer solutions consist of nearly equal concentrations of a weak acid and a salt containing its conjugate base such as acetic acid, $HC_2H_3O_2$, and its salt $NaC_2H_3O_2$.
- The $[H_3O^+]$ is calculated by solving the K_a expression for $[H_3O^+]$, then substituting the values of $[H_3O^+]$, [HA], and K_a into the equation.

$$K_a = \frac{[H_3O^+][C_2H_3O_2^-]}{[HC_2H_3O_2]}$$

Solving for $[H_3O^+]$ gives:

$$[H_3O^+] = K_a \times \frac{[HC_2H_3O_2]}{[C_2H_3O_2^-]} \begin{array}{l} \longleftarrow \text{Weak acid} \\ \longleftarrow \text{Conjugate base} \end{array}$$

- The pH of the buffer is calculated from the $[H_3O^+]$.

$$pH = -\log[H_3O^+]$$

Example: What is the pH of a buffer prepared with 0.40 M $HC_2H_3O_2$ and 0.20 M $C_2H_3O_2^-$ if the K_a of acetic acid is 1.8×10^{-5}?

Answer:

$$[H_3O^+] = K_a \times \frac{[HC_2H_3O_2]}{[C_2H_3O_2^-]} = 1.8 \times 10^{-5} \times \frac{[0.40]}{[0.20]}$$

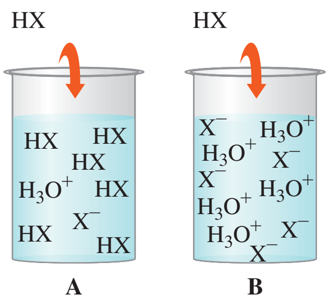
$$= 3.6 \times 10^{-5}\,\text{M}$$

$$pH = -\log[3.6 \times 10^{-5}] = 4.44$$

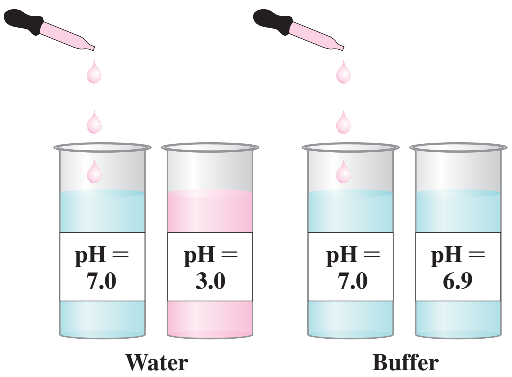
UNDERSTANDING THE CONCEPTS

The chapter sections to review are shown in parentheses at the end of each question.

14.89 Determine if each of the following diagrams represents a strong acid or a weak acid. The acid has the formula HX. (14.3)



14.90 Adding a few drops of a strong acid to water will lower the pH appreciably. However, adding the same number of drops to a buffer does not appreciably alter the pH. Why? (14.9)



14.91 Identify each of the following as an acid or a base: (14.1)
a. H_2SO_4 **b.** RbOH
c. $Ca(OH)_2$ **d.** HI

14.92 Identify each of the following as an acid or a base: (14.1)
a. $Sr(OH)_2$ **b.** H_2SO_3
c. $HC_2H_3O_2$ **d.** CsOH

14.93 Complete the following table: (14.2)

Acid	Conjugate Base
H_2O	
	CN^-
HNO_2	
	$H_2PO_4^-$

14.94 Complete the following table: (14.2)

Base	Conjugate Acid
	HS^-
	H_3O^+
NH_3	
HCO_3^-	

Applications

14.95 Sometimes, during stress or trauma, a person can start to hyperventilate. Then the person might breathe into a paper bag to avoid fainting. (14.9)
a. What changes occur in the blood pH during hyperventilation?
b. How does breathing into a paper bag help return blood pH to normal?

Breathing into a paper bag can help a person who is hyperventilating.

14.96 In the blood plasma, pH is maintained by the carbonic acid–bicarbonate buffer system. (14.9)
a. How is pH maintained when acid is added to the buffer system?
b. How is pH maintained when base is added to the buffer system?

14.97 State whether each of the following solutions is acidic, basic, or neutral: (14.6)
a. sweat, pH 5.2 **b.** tears, pH 7.5
c. bile, pH 8.1 **d.** stomach acid, pH 2.5

14.98 State whether each of the following solutions is acidic, basic, or neutral: (14.6)
a. saliva, pH 6.8 **b.** urine, pH 5.9
c. pancreatic juice, pH 8.0 **d.** blood, pH 7.45

ADDITIONAL QUESTIONS AND PROBLEMS

14.99 Identify each of the following as an acid, base, or salt, and give its name: (14.1)
a. $HBrO_2$ **b.** CsOH
c. $Mg(NO_3)_2$ **d.** $HClO_4$

14.100 Identify each of the following as an acid, base, or salt, and give its name: (14.1)
a. HNO_2 **b.** $MgBr_2$
c. NH_3 **d.** Li_2SO_3

14.101 Complete the following table: (14.2)

Acid	Conjugate Base
HI	
	Cl^-
NH_4^+	
	HS^-

14.102 Complete the following table: (14.2)

Base	Conjugate Acid
F^-	
	$HC_2H_3O_2$
	HSO_3^-
ClO^-	

14.103 Using Table 14.3, identify the stronger acid in each of the following pairs: (14.3)
a. HF or HCN **b.** H_3O^+ or H_2S
c. HNO_2 or $HC_2H_3O_2$ **d.** H_2O or HCO_3^-

14.104 Using Table 14.3, identify the stronger base in each of the following pairs: (14.3)
a. H_2O or Cl^- **b.** OH^- or NH_3
c. SO_4^{2-} or NO_2^- **d.** CO_3^{2-} or H_2O

14.105 Determine the pH and pOH for each of the following solutions: (14.6)
a. $[H_3O^+] = 2.0 \times 10^{-8}$ M
b. $[H_3O^+] = 5.0 \times 10^{-2}$ M
c. $[OH^-] = 3.5 \times 10^{-4}$ M
d. $[OH^-] = 0.0054$ M

14.106 Determine the pH and pOH for each of the following solutions: (14.6)
a. $[OH^-] = 1.0 \times 10^{-7}$ M
b. $[H_3O^+] = 4.2 \times 10^{-3}$ M
c. $[H_3O^+] = 0.0001$ M
d. $[OH^-] = 8.5 \times 10^{-9}$ M

14.107 Are the solutions in problem 14.105 acidic, basic, or neutral? (14.6)

14.108 Are the solutions in problem 14.106 acidic, basic, or neutral? (14.6)

14.109 Calculate the $[H_3O^+]$ and $[OH^-]$ for a solution with each of the following pH values: (14.6)
a. 3.00 **b.** 6.2
c. 8.85 **d.** 11.00

14.110 Calculate the $[H_3O^+]$ and $[OH^-]$ for a solution with each of the following pH values: (14.6)
a. 10.00 **b.** 5.0
c. 6.54 **d.** 1.82

14.111 Solution A has a pH of 4.5, and solution B has a pH of 6.7. (14.6)
- **a.** Which solution is more acidic?
- **b.** What is the $[H_3O^+]$ in each?
- **c.** What is the $[OH^-]$ in each?

14.112 Solution X has a pH of 9.5, and solution Y has a pH of 7.5. (14.6)
- **a.** Which solution is more acidic?
- **b.** What is the $[H_3O^+]$ in each?
- **c.** What is the $[OH^-]$ in each?

14.113 What is the pH and pOH of a solution prepared by dissolving 2.5 g of HCl in water to make 425 mL of solution? (14.6)

14.114 What is the pH and pOH of a solution prepared by dissolving 1.0 g of $Ca(OH)_2$ in water to make 875 mL of solution? (14.6)

CHALLENGE QUESTIONS

The following groups of questions are related to the topics in this chapter. However, they do not all follow the chapter order, and they require you to combine concepts and skills from several sections. These questions will help you increase your critical thinking skills and prepare for your next exam.

14.115 For each of the following: (14.2, 14.3)
1. H_2S **2.** H_3PO_4
- **a.** Write the formula for the conjugate base.
- **b.** Write the K_a expression.
- **c.** Which is the weaker acid?

14.116 For each of the following: (14.2, 14.3)
1. HCO_3^- **2.** $HC_2H_3O_2$
- **a.** Write the formula for the conjugate base.
- **b.** Write the K_a expression.
- **c.** Which is the stronger acid?

14.117 Using Table 14.3, identify the conjugate acid–base pairs in each of the following equations and whether the equilibrium mixture contains mostly products or mostly reactants: (14.2, 14.3)
- **a.** $NH_3(aq) + HNO_3(aq) \rightleftharpoons NH_4^+(aq) + NO_3^-(aq)$
- **b.** $H_2O(l) + HBr(aq) \rightleftharpoons H_3O^+(aq) + Br^-(aq)$

14.118 Using Table 14.3, identify the conjugate acid–base pairs in each of the following equations and whether the equilibrium mixture contains mostly products or mostly reactants: (14.2, 14.3)
- **a.** $HNO_2(aq) + HS^-(aq) \rightleftharpoons H_2S(g) + NO_3^-(aq)$
- **b.** $Cl^-(aq) + H_2O(l) \rightleftharpoons OH^-(aq) + HCl(aq)$

14.119 Complete and balance each of the following: (14.7)
- **a.** $ZnCO_3(s) + H_2SO_4(aq) \longrightarrow$
- **b.** $Al(s) + HNO_3(aq) \longrightarrow$

14.120 Complete and balance each of the following: (14.7)
- **a.** $H_3PO_4(aq) + Ca(OH)_2(s) \longrightarrow$
- **b.** $KHCO_3(s) + HNO_3(aq) \longrightarrow$

14.121 Determine each of the following for a 0.050 M KOH solution: (14.6, 14.7, 14.8)
- **a.** $[H_3O^+]$
- **b.** pH
- **c.** pOH
- **d.** the balanced equation for the reaction with H_2SO_4
- **e.** milliliters of KOH solution required to neutralize 40.0 mL of a 0.035 M H_2SO_4 solution

14.122 Determine each of the following for a 0.10 M HBr solution: (14.6, 14.7, 14.8)
- **a.** $[H_3O^+]$
- **b.** pH
- **c.** pOH
- **d.** the balanced equation for the reaction with LiOH
- **e.** milliliters of HBr solution required to neutralize 36.0 mL of a 0.25 M LiOH solution

14.123 One of the most acidic lakes in the United States is Little Echo Pond in the Adirondacks in New York. Recently, this lake had a pH of 4.2, well below the recommended pH of 6.5. (14.6, 14.8)
- **a.** What are the $[H_3O^+]$ and $[OH^-]$ of Little Echo Pond?
- **b.** What are the $[H_3O^+]$ and $[OH^-]$ of a lake that has a pH of 6.5?
- **c.** One way to raise the pH of an acidic lake (and restore aquatic life) is to add limestone ($CaCO_3$). How many grams of $CaCO_3$ are needed to neutralize 1.0 kL of the acidic water from the lake if the acid is sulfuric acid?

$$H_2SO_4(aq) + CaCO_3(s) \longrightarrow CO_2(g) + H_2O(l) + CaSO_4(aq)$$

A helicopter drops calcium carbonate on an acidic lake to increase its pH.

Applications

14.124 The daily output of stomach acid (gastric juice) is 1000 mL to 2000 mL. Prior to a meal, stomach acid (HCl) typically has a pH of 1.42. (14.6, 14.7, 14.8)
- **a.** What is the $[H_3O^+]$ of stomach acid?
- **b.** One chewable tablet of the antacid Maalox contains 600. mg of $CaCO_3$. Write the neutralization equation, and calculate the milliliters of stomach acid neutralized by two tablets of Maalox.
- **c.** The antacid milk of magnesia contains 400. mg of $Mg(OH)_2$ per teaspoon. Write the neutralization equation, and calculate the number of milliliters of stomach acid that are neutralized by 1 tablespoon of milk of magnesia.

14.125 Calculate the volume, in milliliters, of a 0.150 M NaOH solution that will completely neutralize each of the following: (14.8)
- **a.** 25.0 mL of a 0.288 M HCl solution
- **b.** 10.0 mL of a 0.560 M H_2SO_4 solution

14.126 Calculate the volume, in milliliters, of a 0.215 M NaOH solution that will completely neutralize each of the following: (14.8)

a. 3.80 mL of a 1.25 M HNO_3 solution

b. 8.50 mL of a 0.825 M H_3PO_4 solution

14.127 A solution of 0.205 M NaOH is used to 20.0 mL of a H_2SO_4 solution. If 45.6 mL of the NaOH solution is required to reach the endpoint, what is the molarity of the H_2SO_4 solution? (14.8)

$$H_2SO_4(aq) + 2NaOH(aq) \longrightarrow 2H_2O(l) + Na_2SO_4(aq)$$

14.128 A 10.0-mL sample of vinegar, which is an aqueous solution of acetic acid, $HC_2H_3O_2$, requires 16.5 mL of a 0.500 M NaOH solution to reach the endpoint in a titration. What is the molarity of the acetic acid solution? (14.8)

$$HC_2H_3O_2(aq) + NaOH(aq) \longrightarrow H_2O(l) + NaC_2H_3O_2(aq)$$

14.129 A buffer solution is made by dissolving H_3PO_4 and NaH_2PO_4 in water. (14.9)

a. Write an equation that shows how this buffer neutralizes added acid.

b. Write an equation that shows how this buffer neutralizes added base.

c. Calculate the pH of this buffer if it contains 0.50 M H_3PO_4 and 0.20 M $H_2PO_4^-$. The K_a for H_3PO_4 is 7.5×10^{-3}.

14.130 A buffer solution is made by dissolving $HC_2H_3O_2$ and $NaC_2H_3O_2$ in water. (14.9)

a. Write an equation that shows how this buffer neutralizes added acid.

b. Write an equation that shows how this buffer neutralizes added base.

c. Calculate the pH of this buffer if it contains 0.20 M $HC_2H_3O_2$ and 0.40 M $C_2H_3O_2^-$. The K_a for $HC_2H_3O_2$ is 1.8×10^{-5}.

ANSWERS

Answers to Selected Questions and Problems

14.1 **a.** acid **b.** acid **c.** acid **d.** base **e.** both

14.3 **a.** hydrochloric acid **b.** calcium hydroxide **c.** perchloric acid **d.** nitric acid **e.** sulfurous acid **f.** bromous acid

14.5 **a.** RbOH **b.** HF **c.** H_3PO_4 **d.** LiOH **e.** NH_4OH **f.** HIO_4

14.7 **a.** HI is the acid (hydrogen ion donor), and H_2O is the base (hydrogen ion acceptor).

b. H_2O is the acid (hydrogen ion donor), and F^- is the base (hydrogen ion acceptor).

c. H_2S is the acid (hydrogen ion donor), and $CH_3—CH_2—NH_2$ is the base (hydrogen ion acceptor).

14.9 **a.** F^- **b.** OH^- **c.** HPO_3^{2-} **d.** SO_4^{2-} **e.** ClO_2^-

14.11 **a.** HCO_3^- **b.** H_3O^+ **c.** H_3PO_4 **d.** HBr **e.** $HClO_4$

14.13 **a.** The conjugate acid–base pairs are H_2CO_3/HCO_3^- and H_3O^+/H_2O.

b. The conjugate acid–base pairs are NH_4^+/NH_3 and H_3O^+/H_2O.

c. The conjugate acid–base pairs are HCN/CN^- and HNO_2/NO_2^-.

d. The conjugate acid–base pairs are HF/F^- and $HCHO_2/CHO_2^-$.

14.15 $NH_4^+(aq) + H_2O(l) \rightleftarrows NH_3(aq) + H_3O^+(aq)$

14.17 A strong acid is a good hydrogen ion donor, whereas its conjugate base is a poor hydrogen ion acceptor.

14.19 **a.** HBr **b.** HSO_4^- **c.** H_2CO_3

14.21 **a.** HSO_4^- **b.** HF **c.** HCO_3^-

14.23 **a.** reactants **b.** reactants **c.** products

14.25 $NH_4^+(aq) + SO_4^{2-}(aq) \rightleftarrows NH_3(aq) + HSO_4^-(aq)$

The equilibrium mixture contains mostly reactants because NH_4^+ is a weaker acid than HSO_4^-, and SO_4^{2-} is a weaker base than NH_3.

14.27 **a.** true **b.** false **c.** false **d.** true **e.** false

14.29 **a.** H_2SO_3 **b.** HSO_3^- **c.** H_2SO_3 **d.** HS^- **d.** H_2SO_3

14.31 $H_3PO_4(aq) + H_2O(l) \rightleftarrows H_3O^+(aq) + H_2PO_4^-(aq)$

$$K_a = \frac{[H_3O^+][H_2PO_4^-]}{[H_3PO_4]}$$

14.33 In pure water, $[H_3O^+] = [OH^-]$ because one of each is produced every time a hydrogen ion is transferred from one water molecule to another.

14.35 In an acidic solution, the $[H_3O^+]$ is greater than the $[OH^-]$.

14.37 **a.** acidic **b.** basic **c.** basic **d.** acidic

14.39 **a.** 1.0×10^{-5} M **b.** 1.0×10^{-8} M **c.** 5.0×10^{-10} M **d.** 2.5×10^{-2} M

14.41 **a.** 2.5×10^{-13} M **b.** 2.0×10^{-9} M **c.** 5.0×10^{-11} M **d.** 1.3×10^{-6} M

14.43 In a neutral solution, the $[H_3O^+]$ is 1.0×10^{-7} M and the pH is 7.00, which is the negative value of the power of 10.

14.45 **a.** basic **b.** acidic **c.** basic **d.** acidic **e.** acidic **f.** basic

14.47 An increase or decrease of one pH unit changes the $[H_3O^+]$ by a factor of 10. Thus a pH of 3 is 10 times more acidic than a pH of 4.

14.49 **a.** 4.0 **b.** 8.5 **c.** 9.0 **d.** 3.40 **e.** 7.17 **f.** 10.92

14.51

$[H_3O^+]$	$[OH^-]$	pH	pOH	Acidic, Basic, or Neutral?
1.0×10^{-8} M	1.0×10^{-6} M	8.00	6.00	Basic
3.2×10^{-4} M	3.1×10^{-11} M	3.49	10.51	Acidic
2.8×10^{-5} M	3.6×10^{-10} M	4.55	9.45	Acidic
1.0×10^{-12} M	1.0×10^{-2} M	12.00	2.00	Basic

14.53 1.2×10^{-7} M

14.55 **a.** $ZnCO_3(s) + 2HBr(aq) \longrightarrow CO_2(g) + H_2O(l) + ZnBr_2(aq)$
b. $Zn(s) + 2HCl(aq) \longrightarrow H_2(g) + ZnCl_2(aq)$
c. $HCl(aq) + NaHCO_3(s) \longrightarrow CO_2(g) + H_2O(l) + NaCl(aq)$
d. $H_2SO_4(aq) + Mg(OH)_2(s) \longrightarrow 2H_2O(l) + MgSO_4(aq)$

14.57 **a.** $2HCl(aq) + Mg(OH)_2(s) \longrightarrow 2H_2O(l) + MgCl_2(aq)$
b. $H_3PO_4(aq) + 3LiOH(aq) \longrightarrow 3H_2O(l) + Li_3PO_4(aq)$

14.59 **a.** $H_2SO_4(aq) + 2NaOH(aq) \longrightarrow 2H_2O(l) + Na_2SO_4(aq)$
b. $3HCl(aq) + Fe(OH)_3(s) \longrightarrow 3H_2O(l) + FeCl_3(aq)$
c. $H_2CO_3(aq) + Mg(OH)_2(s) \longrightarrow 2H_2O(l) + MgCO_3(s)$

14.61 To a known volume of a formic acid solution, add a few drops of indicator. Place a solution of NaOH of known molarity in a buret. Add base to acid until one drop changes the color of the solution. Use the volume and molarity of NaOH and the volume of a formic acid solution to calculate the concentration of the formic acid in the sample.

14.63 0.829 M HCl solution

14.65 0.124 M H_2SO_4 solution

14.67 16.5 mL

14.69 **b** and **c** are buffer systems. **b** contains the weak acid H_2CO_3 and its salt $NaHCO_3$. **c** contains HF, a weak acid, and its salt KF.

14.71 **a.** 3 **b.** 1 and 2 **c.** 3 **d.** 1

14.73 pH = 3.35

14.75 The pH of the 0.10 M HF/0.10 M NaF buffer is 3.46.
The pH of the 0.060 M HF/0.120 M NaF buffer is 3.76.

14.77 If you breathe fast, CO_2 is expelled and the equilibrium shifts to lower H_3O^+, which raises the pH.

14.79 If large amounts of HCO_3^- are lost, equilibrium shifts to higher H_3O^+, which lowers the pH.

14.81 pH = 3.70

14.83 2.5×10^{-4} M

14.85 $CaCO_3(s) + 2HCl(aq) \longrightarrow CO_2(g) + H_2O(l) + CaCl_2(aq)$

14.87 0.200 g of $CaCO_3$

14.89 **a.** This diagram represents a weak acid; only a few HX molecules separate into H_3O^+ and X^- ions.
b. This diagram represents a strong acid; all the HX molecules separate into H_3O^+ and X^- ions.

14.91 **a.** acid **b.** base **c.** base **d.** acid

14.93

Acid	Conjugate Base
H_2O	OH^-
HCN	CN^-
HNO_2	NO_2^-
H_3PO_4	$H_2PO_4^-$

14.95 **a.** During hyperventilation, a person will lose CO_2 and the blood pH will rise.
b. Breathing into a paper bag will increase the CO_2 concentration and lower the blood pH.

14.97 **a.** acidic **b.** basic **c.** basic **d.** acidic

14.99 **a.** acid, bromous acid
b. base, cesium hydroxide
c. salt, magnesium nitrate
d. acid, perchloric acid

14.101

Acid	Conjugate Base
HI	I^-
HCl	Cl^-
NH_4^+	NH_3
H_2S	HS^-

14.103 **a.** HF **b.** H_3O^+ **c.** HNO_2 **d.** HCO_3^-

14.105 **a.** pH = 7.70; pOH = 6.30
b. pH = 1.30; pOH = 12.70
c. pH = 10.54; pOH = 3.46
d. pH = 11.73; pOH = 2.27

14.107 **a.** basic **b.** acidic **c.** basic **d.** basic

14.109 **a.** $[H_3O^+] = 1.0 \times 10^{-3}$ M; $[OH^-] = 1.0 \times 10^{-11}$ M
b. $[H_3O^+] = 6 \times 10^{-7}$ M; $[OH^-] = 2 \times 10^{-8}$ M
c. $[H_3O^+] = 1.4 \times 10^{-9}$ M; $[OH^-] = 7.1 \times 10^{-6}$ M
d. $[H_3O^+] = 1.0 \times 10^{-11}$ M; $[OH^-] = 1.0 \times 10^{-3}$ M

14.111 **a.** Solution A
b. Solution A $[H_3O^+] = 3 \times 10^{-5}$ M; Solution B $[H_3O^+] = 2 \times 10^{-7}$ M
c. Solution A $[OH^-] = 3 \times 10^{-10}$ M; Solution B $[OH^-] = 5 \times 10^{-8}$ M

14.113 pH = 0.80; pOH = 13.20

14.115 **a.** **1.** HS^-
2. $H_2PO_4^-$
b. **1.** $\dfrac{[H_3O^+][HS^-]}{[H_2S]}$
2. $\dfrac{[H_3O^+][H_2PO_4^-]}{[H_3PO_4]}$
c. H_2S

14.117 **a.** HNO_3/NO_3^- and NH_4^+/NH_3; equilibrium mixture contains mostly products
b. HBr/Br^- and H_3O^+/H_2O; equilibrium mixture contains mostly products

14.119 **a.** $ZnCO_3(s) + H_2SO_4(aq) \longrightarrow CO_2(g) + H_2O(l) + ZnSO_4(aq)$
b. $2Al(s) + 6HNO_3(aq) \longrightarrow 3H_2(g) + 2Al(NO_3)_3(aq)$

14.121 **a.** $[H_3O^+] = 2.0 \times 10^{-13}$ M
b. pH = 12.70
c. pOH = 1.30
d. $2KOH(aq) + H_2SO_4(aq) \longrightarrow 2H_2O(l) + K_2SO_4(aq)$
e. 56 mL of the KOH solution

14.123 **a.** $[H_3O^+] = 6 \times 10^{-5}$ M; $[OH^-] = 2 \times 10^{-10}$ M
b. $[H_3O^+] = 3 \times 10^{-7}$ M; $[OH^-] = 3 \times 10^{-8}$ M
c. 3 g of $CaCO_3$

14.125 **a.** 48.0 mL of NaOH solution
b. 74.7 mL of NaOH solution

14.127 0.234 M H_2SO_4 solution

14.129 **a.** acid:
$H_2PO_4^-(aq) + H_3O^+(aq) \longrightarrow H_3PO_4(aq) + H_2O(l)$
b. base:
$H_3PO_4(aq) + OH^-(aq) \longrightarrow H_2PO_4^-(aq) + H_2O(l)$
c. pH = 1.72

COMBINING IDEAS from CHAPTERS 11 to 14

CI.21 Methane is a major component of purified natural gas used for heating and cooking. When 1.0 mol of methane gas burns with oxygen to produce carbon dioxide and water vapor, 883 kJ of heat is produced. At STP, methane gas has a density of 0.715 g/L. For transport, the natural gas is cooled to −163 °C to form liquefied natural gas (LNG) with a density of 0.45 g/mL. A tank on a ship can hold 7.0 million gallons of LNG. (2.7, 7.2, 7.3, 8.2, 8.3, 9.5, 10.1, 11.6)

An LNG carrier transports liquefied natural gas.

a. Draw the Lewis structure for methane, which has the formula CH_4.
b. What is the mass, in kilograms, of LNG (assume that LNG is all methane) transported in one tank on a ship?
c. What is the volume, in liters, of LNG (methane) from one tank when the LNG (methane) from one tank is converted to methane gas at STP?
d. Write the balanced chemical equation for the combustion of methane and oxygen in a gas burner, including the heat of reaction.

Methane is the fuel burned in a gas cooktop.

e. How many kilograms of oxygen are needed to react with all of the methane in one tank of LNG?
f. How much heat, in kilojoules, is released after burning all of the methane from one tank of LNG?

CI.22 Automobile exhaust is a major cause of air pollution. One pollutant is nitrogen oxide, which forms from nitrogen and oxygen gases in the air at the high temperatures in an automobile engine. Once emitted into the air, nitrogen oxide reacts with oxygen to produce nitrogen dioxide, a reddish brown gas with a sharp, pungent odor that makes up smog. One component of gasoline is octane, C_8H_{18}, which has a density of 0.803 g/mL. In one year, a typical automobile uses 550 gal of gasoline and produces 41 lb of nitrogen oxide. (2.7, 7.2, 7.3, 8.2, 8.3, 9.2, 11.6)

Two gases found in automobile exhaust are carbon dioxide and nitrogen oxide.

a. Write balanced chemical equations for the production of nitrogen oxide and nitrogen dioxide.
b. If all the nitrogen oxide emitted by one automobile is converted to nitrogen dioxide in the atmosphere, how many kilograms of nitrogen dioxide are produced in one year by a single automobile?
c. Write a balanced chemical equation for the combustion of octane.
d. How many moles of C_8H_{18} are present in 15.2 gal of octane?
e. How many liters of CO_2 at STP are produced in one year from the gasoline used by the typical automobile?

CI.23 A mixture of 25.0 g of CS_2 gas and 30.0 g of O_2 gas is placed in 10.0-L container and heated to 125 °C. The products of the reaction are carbon dioxide gas and sulfur dioxide gas. (7.2, 7.3, 8.2, 9.2, 9.3, 11.1, 11.7)

a. Write a balanced chemical equation for the reaction.
b. How many grams of CO_2 are produced?
c. What is the partial pressure, in millimeters of mercury, of the remaining reactant?
d. What is the final pressure, in millimeters of mercury, in the container?

CI.24 In wine-making, glucose ($C_6H_{12}O_6$) from grapes undergoes fermentation in the absence of oxygen to produce ethanol and carbon dioxide. A bottle of vintage port wine has a volume of 750 mL and contains 135 mL of ethanol (C_2H_6O). Ethanol has a density of 0.789 g/mL. In 1.5 lb of grapes, there are 26 g of glucose. (2.7, 7.2, 7.3, 8.2, 9.2, 12.4)

Port is a type of fortified wine that is produced in Portugal.

When the glucose in grapes is fermented, ethanol is produced.

a. Calculate the volume percent (v/v) of ethanol in the port wine.
b. What is the molarity (M) of ethanol in the port wine?
c. Write the balanced chemical equation for the fermentation reaction of sugar in grapes.
d. How many grams of sugar from grapes are required to produce one bottle of port wine?
e. How many bottles of port wine can be produced from 1.0 ton of grapes (1 ton = 2000 lb)?

CI.25 Consider the following reaction at equilibrium:

$$2H_2(g) + S_2(g) \rightleftharpoons 2H_2S(g) + \text{heat}$$

In a 10.0-L container, an equilibrium mixture contains 2.02 g of H_2, 10.3 g of S_2, and 68.2 g of H_2S. (7.2, 7.3, 13.2, 13.3, 13.4, 13.5)

a. What is the numerical value of K_c for this equilibrium mixture?
b. If more H_2 is added to the equilibrium mixture, how will the equilibrium shift?
c. How will the equilibrium shift if the mixture is placed in a 5.00-L container with no change in temperature?
d. If a 5.00-L container has an equilibrium mixture of 0.300 mol of H_2 and 2.50 mol of H_2S, what is the $[S_2]$ if temperature remains constant?

CI.26 A saturated solution of silver hydroxide has a pH of 10.15. (7.2, 7.3, 13.2, 13.6)

a. Write the solubility product expression for silver hydroxide.
b. Calculate the numerical value of K_{sp} for silver hydroxide.
c. How many grams of silver hydroxide will dissolve in 2.0 L of water?

CI.27 A metal M with a mass of 0.420 g completely reacts with 34.8 mL of a 0.520 M HCl solution to form H_2 gas and aqueous MCl_3. (7.2, 7.3, 8.2, 9.2, 11.7)

a. Write a balanced chemical equation for the reaction of the metal M(s) and HCl(aq).
b. What volume, in milliliters, of H_2 at 720. mmHg and 24 °C is produced?
c. How many moles of metal M reacted?
d. Using your results from part **c**, determine the molar mass and name of metal M.
e. Write the balanced chemical equation for the reaction.

When a metal reacts with a strong acid, bubbles of hydrogen gas form.

CI.28 In a teaspoon (5.0 mL) of a liquid antacid, there are 400. mg of $Mg(OH)_2$ and 400. mg of $Al(OH)_3$. A 0.080 M HCl solution, which is similar to stomach acid, is used to neutralize 5.0 mL of the liquid antacid. (12.4, 14.6, 14.7, 14.8)

An antacid neutralizes stomach acid and raises the pH.

a. Write the equation for the neutralization of HCl and $Mg(OH)_2$.
b. Write the equation for the neutralization of HCl and $Al(OH)_3$.
c. What is the pH of the HCl solution?
d. How many milliliters of the HCl solution are needed to neutralize the $Mg(OH)_2$?
e. How many milliliters of the HCl solution are needed to neutralize the $Al(OH)_3$?

CI.29 A KOH solution is prepared by dissolving 8.57 g of KOH in enough water to make 850. mL of KOH solution. (12.4, 14.6, 14.7, 14.8)

a. What is the molarity of the KOH solution?
b. What is the $[H_3O^+]$ and pH of the KOH solution?
c. Write the balanced chemical equation for the neutralization of KOH by H_2SO_4.
d. How many milliliters of a 0.250 M H_2SO_4 solution is required to neutralize 10.0 mL of the KOH solution?

CI.30 A solution of HCl is prepared by diluting 15.0 mL of a 12.0 M HCl solution with enough water to make 750. mL of HCl solution. (12.4, 14.6, 14.7, 14.8)

a. What is the molarity of the HCl solution?
b. What is the $[H_3O^+]$ and pH of the HCl solution?
c. Write the balanced chemical equation for the reaction of HCl and $MgCO_3$.
d. How many milliliters of the diluted HCl solution is required to completely react with 350. mg of $MgCO_3$?

CI.31 A volume of 200.0 mL of a carbonic acid buffer for blood plasma is prepared that contains 0.403 g of $NaHCO_3$ and 0.0149 g of H_2CO_3. At body temperature (37 °C), the K_a of carbonic acid is 7.9×10^{-7}. (7.2, 7.3, 8.4, 14.9)

$$H_2CO_3(aq) + H_2O(l) \rightleftharpoons HCO_3^-(aq) + H_3O^+(aq)$$

a. What is the $[H_2CO_3]$?
b. What is the $[HCO_3^-]$?
c. What is the $[H_3O^+]$?
d. What is the pH of the buffer?
e. Write a balanced chemical equation that shows how this buffer neutralizes added acid.
f. Write a balanced chemical equation that shows how this buffer neutralizes added base.

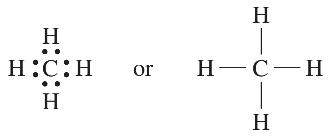
CI.32 In the kidneys, the ammonia buffer system buffers high H_3O^+. Ammonia, which is produced in renal tubules from amino acids, combines with H^+ to be excreted as NH_4Cl. At body temperature (37 °C), the $K_a = 5.6 \times 10^{-10}$. A buffer solution with a volume of 125 mL contains 3.34 g of NH_4Cl and 0.0151 g of NH_3. (7.2, 7.3, 8.4, 14.9)

$$NH_4^+(aq) + H_2O(l) \rightleftharpoons NH_3(aq) + H_3O^+(aq)$$

a. What is the $[NH_4^+]$?
b. What is the $[NH_3]$?
c. What is the $[H_3O^+]$?
d. What is the pH of the buffer?
e. Write a balanced chemical equation that shows how this buffer neutralizes added acid.
f. Write a balanced chemical equation that shows how this buffer neutralizes added base.

ANSWERS

CI.21 a.



b. 1.2×10^7 kg of LNG (methane)

c. 1.7×10^{10} L of LNG (methane)

d. $CH_4(g) + 2O_2(g) \xrightarrow{\Delta} CO_2(g) + 2H_2O(g) + 883\text{ kJ}$

e. 4.8×10^7 kg of O_2

f. 6.6×10^{11} kJ

CI.23 a. $CS_2(g) + 3O_2(g) \xrightarrow{\Delta} CO_2(g) + 2SO_2(g)$

b. 13.8 g of CO_2

c. 37 mmHg

d. 2370 mmHg

CI.25 a. $K_c = 248$

b. If H_2 is added, the equilibrium will shift in the direction of the products.

c. If the volume decreases, the equilibrium will shift in the direction of the products.

d. $[S_2] = 0.280$ M

CI.27 a. $2M(s) + 6HCl(aq) \longrightarrow 3H_2(g) + 2MCl_3(aq)$

b. 233 mL of H_2

c. 6.03×10^{-3} mol of M

d. 69.7 g/mol; gallium

e. $2Ga(s) + 6HCl(aq) \longrightarrow 3H_2(g) + 2GaCl_3(aq)$

CI.29 a. 0.180 M

b. $[H_3O^+] = 5.56 \times 10^{-14}$ M; pH = 13.255

c. $2KOH(aq) + H_2SO_4(aq) \longrightarrow 2H_2O(l) + K_2SO_4(aq)$

d. 3.60 mL

CI.31 a. 0.001 20 M

b. 0.0240 M

c. 4.0×10^{-8} M

d. 7.40

e. $HCO_3^-(aq) + H_3O^+(aq) \longrightarrow H_2CO_3(aq) + H_2O(l)$

f. $H_2CO_3(aq) + OH^-(aq) \longrightarrow HCO_3^-(aq) + H_2O(l)$